Exam 3A
Chem 1142
Spring 2019

Name: KEY

MULTIPLE CHOICE. [2 pts ea.] Record the best response on the scantron sheet. [**40 pts total.**]

Assume all solutions are aqueous and at a temperature of 25 °C, unless stated otherwise.

Q1. Bases turn litmus what color?
A) blue
B) green
C) red
D) yellow

Acids turn litmus red
Bases " " blue

Q2. The Arrhenius definition of an acid is:
A) they donate H^+ ions to other molecules
B) they form H^+ ions in water
C) they accept electron pairs
D) they turn litmus red

Q3. A good example of an Arrhenius base is:
A) NH_3
B) LiOH
C) CH_3CO_2H
D) HCO_3^-

↳ forms OH^- in soln.

$LiOH \longrightarrow Li^+ + OH^-$

Q4. The conjugate acid to HPO_4^{2-} is:
A) H_3PO_4
B) $H_2PO_4^-$
C) PO_4^{3-}
D) H_3O^+

↳ add H^+ to get conj acid

Q5. An example of a strong acid would be:
A) HF
B) HNO_2
C) H_2SO_4
D) H_3PO_4

H_2SO_4 HCl
$HClO_4$ HBr
HNO_3 HI
} six strong acids!

Q6. At 37 °C, K_w is equal to 5.5×10^{-14}. This means the pH of pure water at this temperature will be:
A) 7.00
B) 5.50
C) 13.26
D) 6.63

body temp... neutral
pH = 6.63 !

$K_w = [H^+][OH^-]$

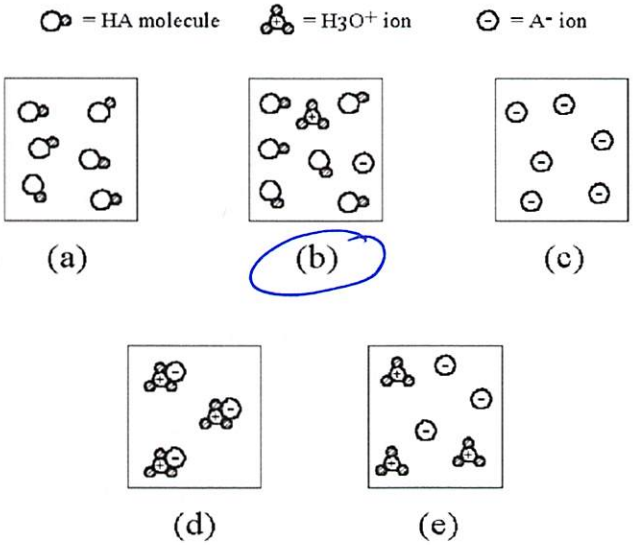
$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$

same conc!

$\Rightarrow 5.5\times10^{-14} = x^2 \quad \Rightarrow x = \sqrt{5.5\times10^{-14}}$

$\Rightarrow x = 2.3\times10^{-7}$, $pH = -\log[H^+] = 6.63$

Q7. Which of the following diagrams represents a snapshot of a very small portion of a beaker containing a weak acid, HA, dissolved in water? Note that the solvent molecules (H_2O) are not shown for clarity.



weak acids partially dissociate!

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

mainly — small amt

largest Ka

Q8. Given three separate solutions containing equal concentrations of formic acid ($K_a = 1.7 \times 10^{-4}$), phenol ($K_a = 1.3 \times 10^{-10}$), and acetic acid ($K_a = 1.8 \times 10^{-5}$), select the response below that has the acids arranged in order of ***increasing*** percent dissociation at equilibrium.

smallest Ka

A) formic < phenol < acetic
B) formic < acetic < phenol
C) acetic < formic < phenol
D) phenol < acetic < formic

$HA + H_2O \rightleftharpoons H_3O^+ + A^-$ (dissociated)

$\Rightarrow$ Ka ↑, % dissoc ↑

Q9. If the concentration of hydroxide ion is 1.8×10^{-4} M in an aqueous solution at 25 °C, what is the pH of the solution?

A) 2.74
B) 3.74
C) 9.26
D) 10.26

$pOH = -\log[OH^-] = 3.74$

@ 25°C, $pH + pOH = 14.00 \Rightarrow pH = 14.00 - pOH = 10.26$

Q10. Which solution has the largest pOH at 25°C: 0.100 M NaOH(aq), 0.100 M $Sr(OH)_2$(aq), or 0.100 M HCl(aq)?

A) 0.100 M NaOH(aq)
B) 0.100 M $Sr(OH)_2$(aq)
C) 0.100 M HCl(aq)
D) Impossible to determine

$pH = -\log[0.100] = 1.000$

$pOH = 14.00 - pH = 13.00$

easiest way...

9 8 7 → pOH
pH ← 5 6 7

$pOH = -\log(0.100) = 1.000$

$pOH = -\log(0.200) = 0.699$

$Sr(OH)_2 \rightarrow Sr^{2+} + 2OH^-$

Big pOH = small pH = strong acid

Q11. What is the pH of 0.25 M LiOH(aq)?

A) 0.60
B) 1.20
C) 10.25
D) 13.40

$LiOH \xrightarrow{100\%} Li^+ + OH^-$

0.25M

$pOH = -\log[OH^-] = 0.60$

$14.00 = pH + pOH \Rightarrow pH = 13.40$

Q12. A Lewis base is a(n):

A) electron-pair donor
B) proton acceptor
C) electron-pair acceptor
D) proton donor

Q13. Which pair of substances will constitute a buffer when found in solution together:

A) $NaNO_2$ / HNO_3
B) KCl / HF
C) HF / NaF
D) $NaNO_2$ / KNO_2

weak acid, HF
conj. base, F^-
counter ion
($Na^+ F^-$)

Q14. A solution containing 0.30 M HClO(aq) ($K_a = 3.5 \times 10^{-8}$) and 0.25 M of NaClO(aq) would have a pH of:

A) 7.46
B) 7.38
C) 7.54
D) 7.31

weak acid
$pK_a = -\log(K_a) = 7.46$
conj. base ClO^-
H-H eq: $pH = pK_a + \log \frac{b}{a} = 7.46 + \log \frac{0.25M}{0.30M} = 7.38$

Q15. A weak monoprotic acid (HA) has a pK_a of 3.94. If we need to prepare a buffer with a pH of 3.74, then we can say for sure that:

A) $[HA] > [A^-]$
B) $[HA] = [A^-] - 0.20$ M
C) $[HA] < [A^-]$
D) $[HA] = [A^-] + 0.20$ M

$pH < pK_a$! So need more HA, $[HA] > [A^-]$
or: $pH = pK_a + \log \frac{b}{a} \Rightarrow pH - pK_a = -0.20 = \log \frac{b}{a}$
$\Rightarrow \frac{[b]}{[a]} = 10^{-0.20} = 0.63$
$\Rightarrow [a] = \frac{1}{0.63} \times [b]$
$[a] = 1.58[b]$
$[a] > [b]$!

Q16. Which of the following acids would it be best to use to prepare a buffer with a pH of 4.25?

A) $HClO_2$, $K_a = 1.1 \times 10^{-2}$ $pK_a = 1.96$
B) HNO_2, $K_a = 4.0 \times 10^{-4}$ $pK_a = 3.40$
C) HCH_2O, $K_a = 1.8 \times 10^{-4}$ $= 3.74$ * closest to pH!
D) H_2CO_3, $K_a = 4.3 \times 10^{-7}$ $= 6.37$

(best buffer is when $pK_a \approx pH$ because then $[b] \approx [a]$)

Q17. Which chemical equation best corresponds to the K_{sp} reaction for calcium carbonate?

A) $CaCO_3(s) \rightleftharpoons Ca^+(aq) + C^{3+}(aq) + 2O^{2-}(g)$
B) $CaCO_3(s) \rightleftharpoons Ca(aq) + CO_3(aq)$
C) $CaCO_3(s) \rightleftharpoons Ca^+(aq) + CO_3^-(aq)$
D) $CaCO_3(s) \rightleftharpoons Ca^{2+}(aq) + CO_3^{2-}(aq)$

Ksp rxn
solid $\rightleftharpoons$ dissolved ions

Q18. Which substance will have the **smallest** molar solubility: $BaSO_4$ ($K_{sp} = 1.07 \times 10^{-10}$), $CaSO_4$ ($K_{sp} = 7.10 \times 10^{-5}$), or FeS ($K_{sp} = 3.72 \times 10^{-19}$).

A) $BaSO_4$
B) $CaSO_4$
C) FeS
D) Impossible to determine

K_{sp} ↓, so fewer dissolved ions
$\Rightarrow$ lower molar sol'y

Q19. In which solution would ammonium fluoride be the **most soluble**?

A) 0.25 M NH_4Br(aq) } NH_4^+ common ion
B) 0.10 M NH_4Cl(aq) } NH_4^+ common ion
C) 0.50 M LiF(aq) } F^- common ion
D) 0.35 M KBr(aq) } no common ion! Highest sol'y!

$NH_4F(s) \rightleftharpoons NH_4^+(aq) + F^-(aq)$
if these [ ]'s are increased, see shift to LHS (Le Châtelier) $\Rightarrow$ lower sol'y!

Q20. The solubility product constant for magnesium hydroxide, $Mg(OH)_2$ is 2.06×10^{-13}. Which solution would form a precipitate immediately upon mixing?

A) A solution containing 1.0×10^{-6} M Mg^{2+} and 1.0×10^{-3} M OH^-
B) A solution containing 1.0×10^{-5} M Mg^{2+} and 1.0×10^{-4} M OH^-
C) A solution containing 1.0×10^{-4} M Mg^{2+} and 1.0×10^{-5} M OH^-
D) A solution containing 1.0×10^{-3} M Mg^{2+} and 1.0×10^{-6} M OH^-

$Mg(OH)_2(s) \rightleftharpoons Mg^{2+}(aq) + 2F^-(aq)$
$Q_{sp} = [Mg^{2+}][F^-]^2$

$Q_{sp} = (10^{-6})(10^{-3})^2 = 10^{-12} > 2.06 \times 10^{-13}$!
so will ppt upon mixing

if $Q_{sp} > K_{sp}$, will shift to LHS + form ppt.

K_{sp} | ← | Q_{sp} →

Short Response.

Show ALL work to receive credit.

Q21. [15 pts.] (a) The pH of 0.13 M HF(aq) is 2.01. Use this information to determine K_a for HF. Be sure to write an ICE chart as part of your answer.

$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$

I	0.13	—	≈0	0
C	$-x$	—	$+x$	$+x$
E	$(0.13-x)$	—	(x)	(x)

$$K_a = \frac{[H_3O^+][F^-]}{[HF]_{eq}} = \frac{x^2}{0.13-x}$$

$$pH = 2.01 \Rightarrow [H_3O^+] = 10^{-pH} = 10^{-2.01} = x \Rightarrow K_a = \frac{(10^{-2.01})^2}{0.13-10^{-2.01}} = 7.9\times10^{-4}$$

(b) Using your calculated K_a, predict the pH of 1.3 M HF(aq).

$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$

I	1.3	—	≈0	0
C	$-x$	—	$+x$	$+x$
E	$(1.3-x)$	—	(x)	(x)

$$K_a = \frac{x^2}{1.3-x} = 7.9\times10^{-4}. \quad \text{Assume } x \ll 1.3$$

$$\Rightarrow 7.9\times10^{-4} \approx \frac{x^2}{1.3} \Rightarrow x = \sqrt{1.3\times7.9\times10^{-4}} = 0.032$$

5% rule? % ionization $= \frac{[H_3O^+]_{eq}}{[HA]_0}\times100 = \frac{0.032}{1.3}\times100 = 2.5\%$ ✓

$pH = -\log[H^+] = -\log[0.032] = 1.49$

(c) What is the percent dissociation of HF in part (a) vs. part (b)?

Part (a) 7.5 % **Part (b)** 2.5 %

% dissoc. $= \frac{x}{0.13}\times100 = \frac{10^{-2.01}}{0.13}\times100 = 7.5\%$

Q22. [15 pts.] **(Be sure to write all relevant chemical equations and ICE/ICF charts for this question.)** 250.0 mL of 0.400 M $HNO_2(aq)$, $K_a = 4.0 \times 10^{-4}$, is mixed with 50.0 mL of 0.800 M NaOH(aq).

(a) Calculate the pH of the solution formed.

$$\#\text{ mol } HNO_2: \quad 250.0\text{mL} \times \frac{1\text{L}}{1000\text{mL}} \times \frac{0.400\text{mol } HNO_2}{1\text{L}} = 0.100\text{ mol } HNO_2$$

$$pK_a = -\log(K_a) = 3.40$$

$$\#\text{mol NaOH}: \quad 50.0\text{mL} \times \frac{1\text{L}}{1000\text{mL}} \times \frac{0.800\text{mol NaOH}}{1\text{L}} = 0.0400\text{ mol NaOH}$$

Rxn: $HNO_2(aq)$ (XS) + $NaOH(aq)$ (LR) $\longrightarrow$ $NaNO_2(aq) + H_2O(l)$

	HNO_2	NaOH	$NaNO_2$	H_2O
I	0.100	0.0400	0	—
C	−0.0400	−0.0400	+0.0400	—
F	0.060	0	0.0400	—

Buffer! (weak acid/conj. base)

$$pH = pK_a + \log\frac{[b]}{[a]} = 3.40 + \log\frac{0.0400\text{mol}/0.3000\text{L}}{0.060\text{ mol}/0.3000\text{L}} = \boxed{3.22}$$

(cancels out!)

(b) If 0.010 mol HCl is added to the solution formed in part (a), what will the final pH be?

Rxn: $HCl(aq)$ (LR) + $NaNO_2(aq)$ (XS) $\longrightarrow$ $HNO_2(aq) + H_2O(l)$

	HCl	$NaNO_2$	HNO_2	H_2O
I	0.010 mol	0.0400	0.060	—
C	−0.010	−0.010	+0.010	—
F	0	0.030	0.070	—

Buffer!

$$pH = pK_a + \log\frac{[\text{base}]}{[\text{acid}]} = 3.40 + \log\frac{0.030}{0.070} = \boxed{3.03}$$

Q23. [15 pts.] Calculate the molar solubility of aluminum hydroxide in pure water vs. in 0.200 M KOH(aq). Comment on the difference. K_{sp} for $Al(OH)_3$ is 2.3×10^{-8}.

Be sure to write the K_{sp} chemical equation and an ICE chart as part of your answer! Explain any assumptions you are making.

in water

K_{sp} rxn: $Al(OH)_3(s) \rightleftharpoons Al^{3+}(aq) + 3OH^-(aq)$

		$Al^{3+}(aq)$	$3OH^-(aq)$
I	___	0	≈ 0
C	"-s"	+s	+3s
E	___	(s)	(3s)

$K_{sp} = [Al^{3+}][OH^-]^3$

$\Rightarrow 2.3\times10^{-8} = (s)(3s)^3 = 27s^4$

$\Rightarrow s = \sqrt[4]{\frac{2.3\times10^{-8}}{27}} = \boxed{5.4\times10^{-3}\,M}$

when dissolving in 0.200M KOH, there's a common ion, OH^-!

$KOH \longrightarrow K^+ + OH^-$

0.200M 0.200M

common ion!

$Al(OH)_3(s) \rightleftharpoons Al^{3+}(aq) + 3OH^-(aq)$

		$Al^{3+}(aq)$	$3OH^-(aq)$
I	___	0	0.200
C	"-s"	+s	+3s
E	___	(s)	(0.200+3s)

$K_{sp} = [Al^{3+}][OH^-]^3$ assuming $3s \ll 0.200$

$\Rightarrow 2.3\times10^{-8} = (s)(0.200+3s)^3 \approx (s)(0.200)^3$

$\Rightarrow s \approx \frac{2.3\times10^{-8}}{(0.200)^3} = \boxed{2.9\times10^{-6}\,M}$

(so, yes! $3s \ll 0.200$)

Comment

molar sol. is much lower when OH^- is already present.

This is expected due to Le Châtelier!

Inc. $[OH^-]$, causes shift to LHS... so less ppt(s) will dissolve!

Q24. [15 pts.] (a) Write out the chemical equation (reaction) corresponding to K_b for $CH_3CH_2NH_2(aq)$, ethylamine.

H^+

$$CH_3CH_2NH_2(aq) + H_2O(l) \rightleftharpoons CH_3CH_2NH_3^+(aq) + OH^-(aq)$$

(b) Write out the chemical equation (reaction) corresponding to K_w.

$$2H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq) .$$

OR

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

(c) Without performing a calculation, explain how the pH of 0.100 M NH_3(aq) compares with 0.100 M LiOH(aq).

LiOH is a strong base (100% dissoc.), so will have a greater pH than NH_3, a weak base of same conc.

(d) Calculate the pH of 12.0 M HCl(aq) (concentrated hydrochloric acid).

— strong acid!

$$HCl \longrightarrow H^+ + Cl^-$$

or $H_2O(l) + HCl(aq) \xrightarrow{100\%} H_3O^+(aq) + Cl^-(aq)$

12.0 M 12.0 M

$$pH = -\log [H^+] = -\log [12.0] = -1.079$$

(3sf.) (3dp)

Bonus question

Predict whether the following aqueous salts will be acidic, basic, neutral, or whether there is not enough information to decide.

NaF basic $Al(NO_3)_3$ acidic LiCl neutral

Na^+: neutral (from strong base, NaOH)

F^-: basic (from weak acid, HF)

Useful Information

Periodic Table of the Elements

IA	IIA											IIIA	IVA	VA	VIA	VIIA	VIIIA
1																	18
1 H 1.008	2											13	14	15	16	17	2 He 4.003
3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.31	3	4	5	6	7	8	9	10	11	12	13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.61	33 As 74.92160	34 Se 78.96	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc [98]	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.60	53 I 126.9	54 Xe 131.3
55 Cs 132.9	56 Ba* 137.3	71 Lu 175.0	72 Hf 178.5	73 Ta 180.9	74 W 183.8	75 Re 186.2	76 Os 190.2	77 Ir 192.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.6	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po [210]	85 At [210]	86 Rn [222]
87 Fr [223]	88 Ra** [226]	103 Lr [262]	104 Rf [261]	105 Db [262]	106 Sg [266]	107 Bh [264]	108 Hs [265]	109 Mt [268]	110 [269]	111 [272]	112 [277]	113	114 [285]	115	116 [289]	117	118 [293]

*	57 La 138.9	58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm [145]	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.50	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0
**	89 Ac [227]	90 Th 232.0	91 Pa 231.0	92 U 238.0	93 Np [237]	94 Pu [244]	95 Am [243]	96 Cm [247]	97 Bk [247]	98 Cf [251]	99 Es [252]	100 Fm [257]	101 Md [258]	102 No [259]

$N_A = 6.022 \times 10^{23}\ \text{mol}^{-1}$

$$R = 8.3145\frac{\text{J}}{\text{mol}\cdot\text{K}} = 0.08206\frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}} \qquad M_1V_1 = M_2V_2$$

$\text{pH} = -\log_{10}[\text{H}^+]$ $\qquad$ $\text{pH} + \text{pOH} = 14.00\ (25\ ^\circ\text{C})$

$K_w = 1.0 \times 10^{-14}\ (25\ ^\circ\text{C})$ $\qquad$ $K_a \cdot K_b = K_w$

$$\text{pH} = \text{p}K_a + \log\frac{[\text{Base}]}{[\text{Acid}]}$$

Given: $ax^2 + bx + c = 0$, then $x = \dfrac{-b \pm \sqrt{b^2 - 4ac}}{2a}$