

NUMERICAL ANSWERS TO ASSIGNED TUTORIAL PROBLEM SETS FOR CHEM206
 FROM KOTZ & TREICHEL'S CHEMISTRY & CHEMICAL REACTIVITY, 6th Ed.

NOTE: the answers from Ch.17 have not been verified. Please report any errors.

Ch.	Q#	Comments
17	2a	NH ₄ ⁺ ammonium ion
	2b	H ₂ CO ₃ carbonic acid
	2c	HBr hydrobromic acid
17	4a	HClO ₄ + H ₂ O → H ₃ O ⁺ + ClO ₄ ⁻ acid A base B conj.acid of B conj.base of A
	4b	NH ₄ ⁺ + H ₂ O → NH ₃ + H ₃ O ⁺ acid A base B conj.base of A conj.acid of B
	4c	HCO ₃ ⁻ + H ₂ O → CO ₃ ²⁻ + H ₃ O ⁺ acid A base B conj.base of A conj.acid of B
17	8	Brønsted acid Brønsted base conjugate base conjugate acid
	8a	CH ₃ CO ₂ H C ₅ H ₅ N CH ₃ CO ₂ ⁻ C ₅ H ₅ NH ⁺
	8b	HSO ₄ ⁻ N ₂ H ₄ SO ₄ ²⁻ N ₂ H ₅ ⁺
	8c	[Al(H ₂ O) ₆] ³⁺ OH ⁻ [Al(H ₂ O) ₅ (OH)] ²⁺ H ₂ O
17	10	3.3x10 ⁻⁴ M = [OH ⁻] ...the solution is basic (pH >7)
17	16a	HF strongest acid (largest K _a); HPO ₄ ²⁻ weakest (smallest K _a)
	16b	F ⁻
	16c	strongest acid (HF) has weakest conjugate base
	16d	weakest acid (HPO ₄ ²⁻) has strongest conjugate base
17	22a	NaNO ₃ → no effect on pH. Neither ion is acidic or basic.
	22b	NaC ₇ H ₅ O ₂ → basic solution. Na ⁺ has no effect on pH, but C ₇ H ₅ O ₂ ⁻ is conjugate base of weak acid HC ₇ H ₅ O ₂ so deprotonates water to some degree & makes solution basic.
	22c	Na ₂ HPO ₄ → basic solution. Na ⁺ has no effect on pH, but HPO ₄ ²⁻ is conjugate base of weak acid H ₂ PO ₄ so deprotonates water to some degree & makes solution basic.
17	28	acetic acid pK _a = -log(1.8x10 ⁻⁵) = 4.74. Acid with smaller pK _a is stronger (larger K _a). Thus, chloroacetic acid is stronger.
17	32	K _a = 10 ^{-pK_a} = 1.1x10 ⁻⁴ . K _b = K _w /K _a = 8.9x10 ⁻¹¹
17	36a	H ₂ S stronger acid than HCO ₃ ⁻ ; thus, eqm lies to right.
	36b	HSO ₄ ⁻ stronger acid than HCN; thus eqm lies to left
	36c	HSO ₄ ⁻ stronger acid than CH ₃ CO ₂ H; eqm lies to left
17	38a	H ₃ O ⁺ (aq) + OCl ⁻ (aq) ⇌ H ₂ O(l) + HOCl(aq)
	38b	Solution should be acidic because the significant species remaining in the solution upon completion of the reaction is HOCl, a weak acid.

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17	42	Given eqm pH data: $[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 0.011\text{M} = [\text{ClCH}_2\text{Cl}_2]$ sub into K expression: find $K_a = 1 \times 10^{-3}$.
17	46a	$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 8.1 \times 10^{-11}\text{ M}$
	46b	$[\text{OH}] = K_w / [\text{H}_3\text{O}^+] = 1.2 \times 10^{-4}\text{ M}$ $K_b = 1.0 \times 10^{-6}$, calculated using equilibrium treatment... this is a moderately weak base.
17	50	initial $[\text{C}_6\text{H}_5\text{OH}] = 0.0166\text{ M}$; use K_a & ICE table... can use approximation OR quadratic formula to solve for x $x = 1.47 \times 10^{-6}\text{ M} = [\text{C}_6\text{H}_5\text{O}] = [\text{H}_3\text{O}^+] \Rightarrow \text{pH} = 5.83$ (2SF in K_a)
17	52	use K_a & ICE table...but approximation is NOT valid (K is too big) use quadratic formula $\Rightarrow x = 8.4 \times 10^{-3}\text{ M} = [\text{BH}^+] = [\text{OH}]$ thus: eqm $[\text{B}] = 0.15 - x = 0.14\text{ M}$
17	56	$[\text{F}] = [\text{H}_3\text{O}^+] = 10^{-2.3} = 0.0050\text{ M}$ at eqm use K_a & ICE table... solve for x = $[\text{HF}] = 0.035\text{ M}$ thus: initial $[\text{HF}] = 0.035\text{ M} + 0.0050\text{ M} = 0.040\text{ M}$
17	60	Find $K_b = K_w / K_a = 7.7 \times 10^{-10}$ Use ICE table & K_b ...approximation ok... $\Rightarrow x = [\text{CH}_3\text{CH}_2\text{CO}_2\text{H}] = [\text{OH}] = 8.8 \times 10^{-6}\text{ M}$ $\text{pH} = 14 - \text{pOH} = 8.94 \Rightarrow [\text{H}_3\text{O}^+] = 1.1 \times 10^{-9}\text{ M}$
17	64a	$\text{pH} = 7$; reacted equimolar amounts of strong acid & strong base
	64b	$\text{pH} > 7$; predominant ion in soln will be HCO_2^- , a weak base
	64c	$\text{pH} < 7$; predominant ion in soln will be $\text{C}_2\text{O}_4^{2-}$, a weak base
17	66	pH will be determined by the first ionization of the acid use K_{a1} and ICE table for loss of first proton only can use approximation... $\Rightarrow x = [\text{H}_3\text{O}^+] = 1.4 \times 10^{-3}\text{ M} \Rightarrow \text{pH} = 2.86$
17	70a	BCl_3 is a Lewis acid
	70b	H_2NNH_2 is a Lewis base
	70c	Ag^+ is a Lewis acid, NH_3 is a Lewis base
17	74	The ion with the more highly charged metal ion, $\text{V}(\text{H}_2\text{O})_6^{3+}$, should be the stronger acid. The stronger interaction between the metal ion and the coordinated water molecules polarizes the O-H bonds more, which makes the δ^+ H more attractive to nearby free water molecules & easier to lose.
17	76	Ethylenediamine can act as a proton acceptor (Brønsted base) and as an electron pair donor (Lewis base)
17	78a	NH_4^+ might lead to an acidic soln; CO_3^{2-} & S^{2-} might lead to a basic solution in water
	78b	Br^- and ClO_4^- have no effect on the pH of a solution
	78c	S^{2-} is the strongest base
	78d	$\text{CO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HCO}_3^-(\text{aq}) + \text{OH}^-(\text{aq})$ $\text{S}^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{HS}^-(\text{aq}) + \text{OH}^-(\text{aq})$

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17	82a	HCO_3^- is a weaker acid than HSO_4^- , so the equilibrium lies predominantly to the left.
	82b	HSO_4^- is a stronger acid than $\text{CH}_3\text{CO}_2\text{H}$, so the equilibrium lies predominantly to the right.
	82c	$\text{Co}(\text{H}_2\text{O})_6^{2+}$ is a weaker acid than $\text{CH}_3\text{CO}_2\text{H}$, so the equilibrium lies predominantly to the left.
17	84	Calculate $[\text{OH}^-] = 0.013 \text{ M}$; $\text{pOH} = 1.87 \Rightarrow \text{pH} = 12.13$
17	86a	$K_b = K_w / K_a = 4.3 \times 10^{-4}$
	86b	The acid is placed directly below $\text{Ni}(\text{H}_2\text{O})_6^{2+}$. HPO_4^{2-} is a weaker acid than $\text{C}_4\text{H}_9\text{NH}_3^+$. PO_4^{3-} is a stronger base than $\text{C}_4\text{H}_9\text{NH}_2$.
17	92	Find $K_a = 4.8 \times 10^{-3}$ from $\text{p}K_a \Rightarrow$ thus $K_b = 2.1 \times 10^{-12}$ Use K_b and ICE table, & approximation... $\Rightarrow x = [\text{OH}^-] = 4.6 \times 10^{-7} \text{ M} \Rightarrow \text{pH} = 14 - \text{pOH} = 7.66$
17	94i	(d) and (e) are acidic solutions
	94ii	(a), (b) and (f) are basic solutions
	94iii	(d) $\text{CH}_3\text{CO}_2\text{H}$ is a stronger acid than NH_4^+
17	96	use pH to find $[\text{C}_6\text{H}_4\text{NO}_2^+] = [\text{H}_3\text{O}^+] = 2.0 \times 10^{-3} \text{ M}$ and initial $[\text{C}_6\text{H}_5\text{NO}_2] = 0.14 \text{ M}$ calculated from data use ICE table & sub into K_a expression $\Rightarrow K_a = 3.0 \times 10^{-5}$
17	98	pH determined by first ionization of base (polyprotic...) Use K_{b1} & ICE table for accepting 1 st proton...approximation ok $\Rightarrow x = [\text{OH}^-] = 1.2 \times 10^{-4} \text{ M} \Rightarrow \text{pH} = 14 - \text{pOH} = 10.07$
17	102	Desired net rxn: $\text{NH}_3(\text{aq}) + \text{H}_3\text{O}^+(\text{aq}) \rightleftharpoons \text{H}_2\text{O}(\ell) + \text{NH}_4^+(\text{aq})$ Can think of it as adding rxn equations for: ① rxn of NH_3 with H_2O & ② rxn of OH^- with H_3O^+ When add reactions, must multiply equilibrium constants to find net equilibrium constant So, $K_{\text{net}} = (K_{b,\text{NH}_3})(1/K_w) = 1.8 \times 10^9$
17	110	H_2SeO_4 should be the stronger acid. In oxoacids, the acid with more oxygen atoms experiences a stronger inductive effect, and has more charge delocalization (due to larger # of resonance structures) in the conjugate base. Thus, the conjugate base is more stable (less basic), which corresponds to the acid form being stronger. Experiment to show this: measure the pH of a solution of each acid, at same concentration, e.g., 0.1 M. The solution containing the stronger acid will have the lower pH.
17	112	highest concentration: $\text{H}_2\text{O} > \text{H}_2\text{C}_2\text{O}_4 > \text{HC}_2\text{O}_4^- = \text{H}_3\text{O}^+ > \text{C}_2\text{O}_4^{2-} > \text{OH}^-$ lowest concentration.
17	114a	The increasing acidity as Br atoms replace H atoms is due to the inductive effect of the Br atoms (which results from their higher electronegativity).
	114b	The solution containing the strongest acid ($\text{Br}_3\text{CCO}_2\text{H}$) will have the lowest pH, and the solution containing the weakest acid ($\text{CH}_3\text{CO}_2\text{H}$) will have the highest pH.
17	116	Dissolved CO_2 gas (mainly) or dissolved metal ions can cause bottled water to be slightly acidic.