

Equilibrium Problems
CHE152
Sherman Henzel

ICE Tables

To ICE or not to ICE, that is the question. ICE tables are used when initial (I) concentrations are given. If only equilibrium concentrations are given, then ICE tables are not employed. ICE problems include pH determination of weak acids and bases, equilibrium concentrations, K determinations, reacting species such as buffer reactions and salts that can form a precipitate, titrations of weak acids and bases, solubility problems, and precipitation problems.

Examples:

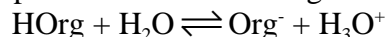
1. Determine the pH of a solution that is 0.1500 M in $\text{HC}_2\text{H}_3\text{O}_2$.

0.1500 M $\text{HC}_2\text{H}_3\text{O}_2$ is an initial concentration and an ICE table is employed.

answer: pH = 2.79

Similar problems: Chapter 18 pages 11, 12, 15, and 17.

2. An organic acid HOrg (MW = 325.00 g/mol) is soluble in water to the extent of 1.58 g HOrg/L. If the pH of a 500.0 mL HOrg solution has a pH of 4.821, what is the K_a for this acid?



The pH is measured at equilibrium, but the solubility of the acid is an initial concentration and an ICE table is employed.

answer: $K_a = 4.70 \times 10^{-8}$

Similar problems: Chapter 18 pages 13 and 14.

3. Calculate the pH of a solution that is 0.1250 M in hydroxylamine.

0.1250 M hydroxylamine is an initial concentration and an ICE table is employed.

answer: pH = 9.53

Similar problems: Chapter 18 page 18.

4. Calculate the pH of a solution that is 0.1150 M in NaOCl.

0.1150 M NaOCl is an initial concentration and an ICE table is employed.

answer: pH = 10.30

Similar problems: Chapter 18 pages 20, 21, 29, 30, and 31.

5. Calculate the $[\text{H}_2\text{S}]_e$, $[\text{HS}^-]_e$, $[\text{S}^{2-}]_e$ and $[\text{H}_3\text{O}^+]_e$ in a solution that is 0.1750 M in H_2S .

0.1750 M H_2S is an initial concentration and an ICE table is employed.

answer: $[\text{H}_2\text{S}] = 0.1750 \text{ M}$, $[\text{HS}^-] = 1.4 \times 10^{-4} \text{ M}$, $[\text{H}_3\text{O}^+] = 1.4 \times 10^{-4} \text{ M}$, $[\text{S}^{2-}] = 1.0 \times 10^{-14}$

Similar problems: Chapter 18 pages 23 and 24.

6. Determine the pH in a solution that is 0.1500 M in HNO_2 and 0.1250 M in HCl .

0.1250 M HNO_2 and 0.1500 M HCl are both initial concentrations and an ICE table is employed.

answer: $\text{pH} = 0.90$

Similar problems: Chapter 19 page 2.

7. Calculate the pH of a solution that is 0.5000 M in $\text{HC}_2\text{H}_3\text{O}_2$ and 0.2500 M in $\text{NaC}_2\text{H}_3\text{O}_2$.

0.5000 M $\text{HC}_2\text{H}_3\text{O}_2$ and 0.2500 M $\text{NaC}_2\text{H}_3\text{O}_2$ are both initial concentrations and an ICE table is employed.

answer: $\text{pH} = 4.46$

Similar problems: Chapter 19 page 3.

8. Calculate the pH of a buffer that is 0.1025 M in HONH_2 and 0.2000 M HONH_3Cl .

0.1025 M HONH_2 and 0.2000 M HONH_3Cl are both initial concentrations and an ICE table is employed.

answer: $\text{pH} = 5.67$

Similar problems: Chapter 19 pages 5 and 8 (bottom).

9. What is the effect of adding 0.01000 moles of NaOH to 1.00 L of the buffer in problem 8 above?

0.01000 M NaOH , a strong base, will react with the conjugate acid of the buffer and an ICE table is employed.

answer: $\Delta\text{pH} = 0.06 \text{ pH units}$

Similar problems: Chapter 19 pages 6, 7, 8, 9 and 10.

10. 50.00 mL of 0.2000 M $\text{HC}_2\text{H}_3\text{O}_2$ are titrated with 0.4000 M NaOH. Calculate the pH after 15.00 mL of NaOH are added.

Weak acid or weak base titrations require an ICE table from the initial addition of titrant until the equivalence point. Post equivalence an ICE table is not required.

answer: pH = 4.94

Similar problems: Chapter 19 pages 21 to 29.

11. Calculate the solubility of AgCl in water.

Solubility problems employ ICE tables.

answer: $S = 1.3 \times 10^{-5} \text{ M}$

Similar problems: Chapter 19 page 38.

12. Calculate the solubility of CaF_2 in a solution that is 0.01500 M in NaF.

Solubility problem (0.01500 M NaF is an initial concentration) employs ICE table.

answer: $S = 2.4 \times 10^{-5} \text{ M}$

Similar problems: Chapter 19 pages 42 and 43.

13. A 15.00 mL sample of 0.01500 M AgNO_3 is added to 20.00 mL of 0.02500 M KI. Will AgI precipitate? Will the precipitation of Ag^+ be complete?

AgNO_3 and KI can react to form a precipitate and ICICE table is employed.

answer: $[\text{Ag}^+] = 1.1 \times 10^{-14} \text{ M}$; ppt is complete

Similar problems: Chapter 19 pages 46 and 47.

14. Calculate the solubility of Ag_2CO_3 in a solution buffered to a pH of 3.00.

Solubility problems employ ICE tables.

answer: $S = 0.029 \text{ M}$

Similar problems: Chapter 19 pages 52 and 53.

15. What is the molar solubility of AgCl in 0.1015 M NH_3 ?

Solubility problems employ ICE tables.

answer: $S = 4.9 \times 10^{-3} \text{ M}$

Similar problems: Chapter 19 pages 60 and 61.

16. What is the solubility of FeS in 1.050 M HCl?

Solubility problems employ ICE tables.

$$\text{answer: } S = 0.5_{21} \text{ M}$$

Similar problems: Chapter 19 pages 66 and 67.

No ICE Tables

ICE tables are not employed when equilibrium (E) concentrations are given. . Non-ICE problems include pH determination of strong, buffer preparation, titrations of strong acids and bases, post equivalence point for weak acid and base titrations, K_{sp} determination, calculation of Q, ion separation by precipitation and preventing precipitation.

17. Calculate the $[H_3O^+]$, $[OH^-]$ and $[Cl^-]$ in 500.0 mL of a solution that is 0.1115 M in HCl.

Strong acid does not employ and ICE table.

$$\text{answer: } [H_3O^+] = [NO_3^-] = 0.1115 \text{ M}, [OH^-] = 9.0 \times 10^{-14} \text{ M}$$

Similar problems: Chapter 18 pages 7 and 8.

18. What mass of $NaNO_2$ must be added to 250.0 mL of 0.350 M HNO_2 in order to produce a buffer with a pH of 3.50?

Buffer preparation involves all concentrations at equilibrium and does not employ ICE tables.

$$\text{answer: } 9.7 \text{ f } NaNO_2$$

Similar problems: Chapter 19 page 4.

19. Describe the preparation of 250.0 mL of a buffer with a pH of 5.000 from 0.1500 M $HC_2H_3O_2$ and 0.2000 M $NaC_2H_3O_2$.

Buffer preparation involves all concentrations at equilibrium and does not employ ICE tables.

$$\text{answer: Mix } 108 \text{ mL of } 0.1500 \text{ M } HC_2H_3O_2 \text{ with } 142 \text{ mL of } 0.2000 \text{ M } NaC_2H_3O_2.$$

Similar problems: Chapter 19 pages 12 and 13.

20. 25.00 mL of 0.1500 M HCl are titrated with 0.3000 M NaOH. Calculate the pH after 7.50 mL of NaOH are added.

Strong acid and strong base titrations do not employ an ICE table.

answer: pH = 1.34

Similar problems: Chapter 19 pages 15 to 20.

21. 50.00 mL of 0.2000 M HC₂H₃O₂ are titrated with 0.4000 M NaOH. Calculate the pH after 30.00 mL of NaOH are added.

Weak acid strong base titrations post equivalence point do not employ ICE tables.

answer: pH = 12.40

Similar problems: Chapter 19 pages 25 and 29.

22. A 125.0 mL sample is removed from a water solution saturated with Ca(OH)₂ at 25.0 °C. The water is completely evaporated from the solution and a deposit of 0.1030 g of Ca(OH)₂ is obtained. Calculate the K_{sp} of Ca(OH)₂ at 25.0 °C.

In a saturated solution all concentrations are at equilibrium and an ICE table is not employed.

answer: K_{sp} = 5.502 x 10⁻⁶

Similar problems: Chapter 19 page 36.

23. It is desired to separate Ba²⁺ from Ag⁺ by the slow addition of SO₄²⁻ to a solution in which the concentration of Ba²⁺ is 0.002500 M and the concentration of Ag⁺ is 0.0001250 M.

- Which ion will precipitate first and at what [SO₄²⁻]?
- What will be the concentration of the first ion to precipitate when the second ion just begins to precipitate?
- Is the separation effective?

Separation problems involve equilibrium concentrations and ICE tables are not employed.

answer: (a) Ba²⁺ will precipitate first when [SO₄²⁻] = 4.4 x 10⁻⁸ M
(b) [Ba²⁺] = 1.2 x 10⁻¹³ M
(c) Yes.

Similar problems: Chapter 19 pages 40 and 41.

24. Eight drops of 0.01500 M NaCl are added to 175.0 mL of 0.007500 M AgNO₃. Will AgCl precipitate? 20 drops equal 1 mL.

Determination of Q does not employ an ICE table.

answer: $Q > K_{sp}$ therefore AgCl (s) will form

Similar problem: Chapter 19 page 44 and 45.

25. What concentration of C₂H₅NH₃⁺ must be maintained in order to prevent the precipitation of Mg(OH)₂ in a solution that is 0.01250 M in Mg(NO₃)₂ and 0.003500 M in C₂H₅NH₂?

Preventing precipitation problems do not require an ICE table.

answer: 0.041 M

Similar problems: Chapter 19 page 51.

26. What is the minimum concentration of NH₃ required to prevent AgCl from precipitating in a solution that is 0.0002500 M in AgNO₃ and 0.00009500 M in NaCl?

Preventing precipitation problems do not require an ICE table.

answer: 0.0034 M

Similar problems: Chapter 19 pages 58 and 59.

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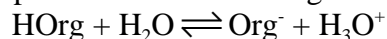
1. Determine the pH of a solution that is 0.1500 M in HC₂H₃O₂.

	HC ₂ H ₃ O ₂	+	H ₂ O	⇌	C ₂ H ₃ O ₂ ⁻	+	H ₃ O ⁺
I	0.1500				0		0
C	-x				+x		+x
E _x	0.1500 - x				x		x
E	≈0.1500						

$$\frac{[\text{C}_2\text{H}_3\text{O}_2^-][\text{H}_3\text{O}^+]}{[\text{HC}_2\text{H}_3\text{O}_2]} = 1.74 \times 10^{-5} = \frac{x^2}{0.1500}; x = 0.00162$$

$$[\text{H}_3\text{O}^+] = 0.00162 \text{ M}; \text{pH} = 2.79$$

2. An organic acid HOrg (MW = 325.00 g/mol) is soluble in water to the extent of 1.58 g HOrg/L. If the pH of a 500.0 mL HOrg solution has a pH of 4.821, what is the K_a for this acid?



$$[\text{Horg}]_i = \frac{1.58 \text{ g HOrg}}{\text{L}} \times \frac{1 \text{ mol HOrg}}{325.00 \text{ g HOrg}} = \frac{0.00486 \text{ mol HOrg}}{\text{L}}$$

$$\text{pH} = 4.821; [\text{H}_3\text{O}^+]_E = 1.51 \times 10^{-5} \text{ M}$$

	HOrg	+	H ₂ O	⇌	Org ⁻	+	H ₃ O ⁺
I	0.00486				0		0
C	-x				+x		+x
E _x	0.00486-x				x		x
E	0.00485				1.51 x 10 ⁻⁵		1.51 x 10 ⁻⁵

$$K_a = \frac{[\text{Org}^-][\text{H}_3\text{O}^+]}{[\text{HOrg}]} = \frac{(1.51 \times 10^{-5})^2}{0.00485} = 4.70 \times 10^{-8}$$

3. Calculate the pH of a solution that is 0.1250 M in hydroxylamine.

	HONH ₂	+	H ₂ O	⇌	HONH ₃ ⁺	+	OH ⁻
I	0.1250				0		0
C	-x				+x		+x
E _x	0.1250 - x				x		x
E	≈0.1250						

$$K_b = \frac{[\text{HONH}_3^+][\text{OH}^-]}{[\text{HONH}_2]} = 9.1 \times 10^{-9} = \frac{x^2}{0.1250}; x = 3.4 \times 10^{-5}$$

$$[\text{OH}^-] = 3.4 \times 10^{-5} \text{ M}; \text{pOH} = 4.47; \text{pH} = 9.53$$

4. Calculate the pH of a solution that is 0.1150 M in NaOCl.

	OCl ⁻	+	H ₂ O	⇌	HOCl	+	OH ⁻
I	0.1150				0		0
C	-x				+x		+x
E _x	0.1150 - x				x		x
E	≈0.1150						

$$K_b = \frac{[\text{HOCl}][\text{OH}^-]}{[\text{OCl}^-]} = \frac{K_w}{K_a} = \frac{1.00 \times 10^{-14}}{2.95 \times 10^{-8}} = 3.39 \times 10^{-7}$$

$$\frac{x^2}{0.1150} = 3.39 \times 10^{-7}; x = 1.97 \times 10^{-4}$$

$$[\text{OH}^-] = 1.97 \times 10^{-4} \text{ M}; \text{pOH} = 3.70; \text{pH} = 10.30$$

5. Calculate the $[H_2S]_e$, $[HS^-]_e$, $[S^{2-}]_e$ and $[H_3O^+]_e$ in a solution that is 0.1750 M in H_2S .

	H_2S	+	H_2O	\rightleftharpoons	HS^-	+	H_3O^+
I	0.1750				0		0
C	-x				+x		+x
E_x	$0.1750 - x$				x		x
E	≈ 0.1750						

$$K_{a1} = \frac{[HS^-][H_3O^+]}{[H_2S]} = 1.1 \times 10^{-7} = \frac{x^2}{0.1750}; x = 1.4 \times 10^{-4}$$

$$[H_2S] = 0.1750 \text{ M}, [HS^-] = 1.4 \times 10^{-4} \text{ M}, [H_3O^+] = 1.4 \times 10^{-4} \text{ M}$$

	HS^-	+	H_2O	\rightleftharpoons	S^{2-}	+	H_3O^+
I	1.4×10^{-4}				0		1.4×10^{-4}
C	-x				+x		+x
E_x	$1.4 \times 10^{-4} - x$				x		$1.4 \times 10^{-4} + x$
E	$\approx 1.4 \times 10^{-4}$						$\approx 1.4 \times 10^{-4}$

$$K_{a2} = \frac{[S^{2-}][H_3O^+]}{[HS^-]} = 1.0 \times 10^{-14} = \frac{x(1.4 \times 10^{-4})}{(1.4 \times 10^{-4})}; x = 1.0 \times 10^{-14}; [S^{2-}] = 1.0 \times 10^{-14} \text{ M}$$

6. Determine the pH in a solution that is 0.1500 M in HNO_2 and 0.1250 M in HCl .

	HNO_2	+	H_2O	\rightleftharpoons	NO_2^-	+	H_3O^+
I	0.1500				0		0.1250
C	-x				+x		+x
E_x	$0.1250 - x$				x		x
E	≈ 0.1500						≈ 0.1250

$$K_a = \frac{[NO_2^-][H_3O^+]}{[HNO_2]} = 5.13 \times 10^{-4} = \frac{0.1250x}{0.1500}; x = 6.16 \times 10^{-4}$$

$$[H_3O^+] = 0.1250 \text{ M}; \text{pH} = 0.90$$

7. Calculate the pH of a solution that is 0.5000 M in $\text{HC}_2\text{H}_3\text{O}_2$ and 0.2500 M in $\text{NaC}_2\text{H}_3\text{O}_2$.

	$\text{HC}_2\text{H}_3\text{O}_2$	+	H_2O	\rightleftharpoons	$\text{C}_2\text{H}_3\text{O}_2^-$	+	H_3O^+
I	0.5000				0.2500		0
C	-x				+x		+x
E_x	$0.5000 - x$				$0.2500 + x$		x
E	≈ 0.5000				≈ 0.2500		

$$\frac{[\text{C}_2\text{H}_3\text{O}_2^-][\text{H}_3\text{O}^+]}{[\text{HC}_2\text{H}_3\text{O}_2]} = 1.74 \times 10^{-5} = \frac{0.2500x}{0.5000}; x = 3.48 \times 10^{-5}$$

$$[\text{H}_3\text{O}^+] = 3.48 \times 10^{-5} \text{ M}; \text{pH} = 4.46$$

8. Calculate the pH of a buffer that is 0.1025 M in HONH_2 and 0.2000 M HONH_3Cl .

	HONH_2	+	H_2O	\rightleftharpoons	HONH_3^+	+	OH^-
I	0.1025				0.2000		0
C	-x				+x		+x
E_x	$0.1025 - x$				$0.2000 + x$		x
E	≈ 0.1025				≈ 0.2000		

$$K_b = \frac{[\text{HONH}_3^+][\text{OH}^-]}{[\text{HONH}_2]} = 9.1 \times 10^{-9} = \frac{0.2000x}{0.1025}; x = 4.7 \times 10^{-9}$$

$$[\text{OH}^-] = 4.7 \times 10^{-9} \text{ M}; \text{pOH} = 8.33; \text{pH} = 5.67$$

9. What is the effect of adding 0.01000 moles of NaOH to 1.00 L of the buffer in problem 8 above?

	HONH_3^+	+	OH^-	\rightleftharpoons	HONH_2	+	H_2O
I	0.2000		0.01000		0.1025		
C	-0.01000		-0.01000		+0.01000		
I	0.1900		0		0.1125		
C	+x		+x		-x		
E_x	$0.1900 + x$		x		$0.1125 - x$		
E	≈ 0.1900				≈ 0.1125		

$$K_b = \frac{[\text{HONH}_3^+][\text{OH}^-]}{[\text{HONH}_2]} = 9.1 \times 10^{-9} = \frac{0.1900x}{0.1125}; x = 5.4 \times 10^{-9}$$

$$[\text{OH}^-] = 5.4 \times 10^{-9} \text{ M}; \text{pOH} = 8.27; \text{pH} = 5.73; \Delta\text{pH} = 5.73 - 5.67 = 0.06 \text{ pH units}$$

10. 50.00 mL of 0.2000 M $\text{HC}_2\text{H}_3\text{O}_2$ are titrated with 0.4000 M NaOH. Calculate the pH after 15.00 mL of NaOH are added.

$$[\text{HC}_2\text{H}_3\text{O}_2] = \frac{0.2000 \text{ M}}{1} \times \frac{50.00 \text{ mL}}{65.00 \text{ mL}} = 0.1538 \text{ M}$$

$$[\text{OH}^-] = \frac{0.4000 \text{ M}}{1} \times \frac{15.00 \text{ mL}}{65.00 \text{ mL}} = 0.09231 \text{ M}$$

	$\text{HC}_2\text{H}_3\text{O}_2$	+	OH^-	\rightleftharpoons	$\text{C}_2\text{H}_3\text{O}_2^-$	+	H_2O
I	0.1538		0.09231		0		
C	-0.09231		-0.09231		+0.09231		
I	0.0615		0		0.09231		
C	+x		+x		-x		
E_x	$0.0615 + x$				$0.09231 - x$		
E	≈ 0.0615				≈ 0.09231		

$$K_b = \frac{[\text{HC}_2\text{H}_3\text{O}_2][\text{OH}^-]}{[\text{C}_2\text{H}_3\text{O}_2^-]} = \frac{K_a}{K_a} = \frac{1.00 \times 10^{-14}}{1.74 \times 10^{-5}} = \frac{0.0615x}{0.09231}; x = 8.63 \times 10^{-10}$$

$$[\text{OH}^-] = 8.63 \times 10^{-10} \text{ M}; \text{pOH} = 9.06; \text{pH} = 4.94$$

11. Calculate the solubility of AgCl in water.

	AgCl (s)	\rightleftharpoons	Ag^+	+	Cl^-
I			0		0
C			+x		+x
E_x			x		x
E					

$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = 1.8 \times 10^{-10} = x^2; x = 1.3 \times 10^{-5}$$

$$S = 1.3 \times 10^{-5} \text{ M}$$

12. Calculate the solubility of CaF_2 in a solution that is 0.01500 M in NaF.

	CaF_2 (s)	\rightleftharpoons	Ca^{+2}	+	2F^-
I			0		0.01500
C			+x		+2x
E _x			x		0.01500 + 2x
E					≈ 0.01500

$$K_{\text{sp}} = [\text{Ca}^{2+}][\text{F}^-]^2 = 5.3 \times 10^{-9} = (0.01500)^2 x; x = 2.4 \times 10^{-5}$$

$$S = 2.4 \times 10^{-5} \text{ M}$$

13. A 15.00 mL sample of 0.01500 M AgNO_3 is added to 20.00 mL of 0.02500 M KI. Will AgI precipitate? Will the precipitation of Ag^+ be complete?

$$\frac{0.01500 \text{ mol AgNO}_3}{\text{L}} \times \frac{1 \text{ mol Ag}^+}{1 \text{ mol AgNO}_3} \times \frac{15.00 \text{ mL}}{35.00 \text{ mL}} = 0.006429 \text{ M Ag}^+$$

$$\frac{0.02500 \text{ mol KI}}{\text{L}} \times \frac{1 \text{ mol I}^-}{1 \text{ mol KI}} \times \frac{20.00 \text{ mL}}{35.00 \text{ mL}} = 0.01429 \text{ M I}^-$$

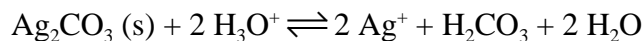
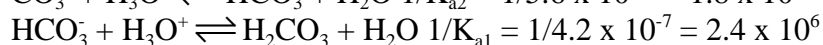
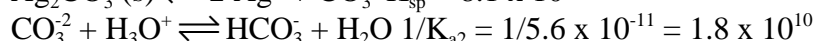
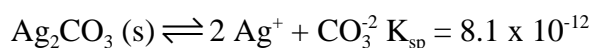
$$Q = [\text{Ag}^+][\text{I}^-] = 0.006429 \times 0.01429 = 9.2 \times 10^{-5}; Q > K_{\text{sp}} \text{ ppt will form}$$

	AgI (s)	\rightleftharpoons	Ag^+	+	I^-
I			0.006429		0.01429
C			-0.006429		-0.006429
I			0		0.00786
C			+x		+x
E _x			x		0.00786 + x
E					≈ 0.00786

$$K_{\text{sp}} = [\text{Ag}^+][\text{I}^-] = 8.5 \times 10^{-17} = 0.00786x; x = 1.1 \times 10^{-14}$$

$$[\text{Ag}^+] = 1.1 \times 10^{-14} \text{ M}; \text{ ppt is complete}$$

14. Calculate the solubility of Ag_2CO_3 in a solution buffered to a pH of 3.00.



$$K = 8.1 \times 10^{-12} \times 1.8 \times 10^{10} \times 2.4 \times 10^6 = 3.5 \times 10^5$$

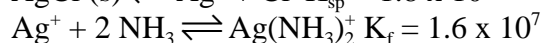
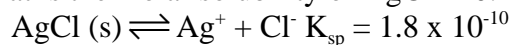
$$\text{pH} = 3.00; [\text{H}_3\text{O}^+] = 1.0 \times 10^{-3} \text{ M}$$

	$\text{Ag}_2\text{CO}_3 (\text{s})$	$+$	$2 \text{H}_3\text{O}^+$	\rightleftharpoons	2Ag^+	$+$	H_2CO_3	$+$	$2 \text{H}_2\text{O}$
I			1.0×10^{-3}		0		0		
C			-----		+2x		+x		
E_x			-----		2x		x		
E			1.0×10^{-3}						

$$K = \frac{[\text{Ag}^+]^2 [\text{H}_2\text{CO}_3]}{[\text{H}_3\text{O}^+]^2} = 3.5 \times 10^5 = \frac{(2x)^2 x}{(1.0 \times 10^{-3})^2} = \frac{4x^3}{1.0 \times 10^{-6}}; x = 0.44$$

$$S = 0.44 \text{ M}$$

15. What is the molar solubility of AgCl in 0.1015 M NH_3 ?



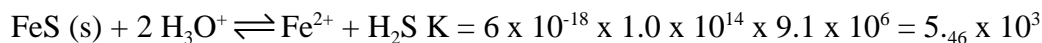
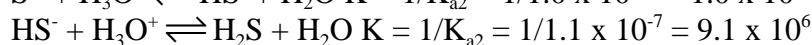
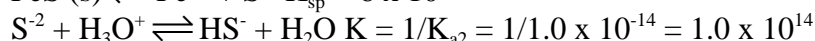
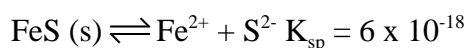
	$\text{AgCl} (\text{s})$	$+$	2NH_3	\rightleftharpoons	$\text{Ag}(\text{NH}_3)_2^+$	$+$	Cl^-
I			0.1015		0		0
C			-2x		+x		+x
E_x			0.1015 - 2x		x		x

$$K = \frac{[\text{Ag}(\text{NH}_3)_2^+][\text{Cl}^-]}{[\text{NH}_3]^2} = 0.0029 = \frac{x^2}{(0.1015 - 2x)^2}$$

$$\sqrt{\frac{x^2}{(0.1015 - 2x)^2}} = \sqrt{0.0029}; \frac{x}{0.1015 - 2x} = 0.053; x = 4.9 \times 10^{-3}$$

$$S = 4.9 \times 10^{-3} \text{ M}$$

16. What is the solubility of FeS in 1.050 M HCl?



	FeS (s) +	2 H ₃ O ⁺	⇌	Fe ²⁺	+	H ₂ S	+	2 H ₂ O
I		1.050		0		0		
C		-2x		+x		+x		
E _x		1.050 - 2x		x		x		

$$K = \frac{[\text{Fe}^{2+}][\text{H}_2\text{S}]}{[\text{H}_3\text{O}^+]^2} = 5.46 \times 10^3 = \frac{x^2}{(1.050 - 2x)^2}$$

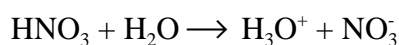
$$\sqrt{5.46 \times 10^3} = \sqrt{\frac{x^2}{(1.050 - 2x)^2}}; 7.39 \times 10^1 = \frac{x}{1.050 - 2x}; x = 0.5_{21}$$

$$S = 0.5_{21} \text{ M}$$

Non-ICE Tables

Examples:

17. Calculate the [H₃O⁺], [OH⁻] and [NO₃⁻] in 500.0 mL of a solution that is 0.1115 M in HNO₃.



$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.00 \times 10^{-14}; 0.1115[\text{OH}^-] = 1.00 \times 10^{-14}; [\text{OH}^-] = 9.0 \times 10^{-14} \text{ M}$$

$$[\text{H}_3\text{O}^+] = [\text{NO}_3^-] = 0.1115 \text{ M}, [\text{OH}^-] = 9.0 \times 10^{-14} \text{ M}$$

18. What mass of NaNO₂ must be added to 250.0 mL of 0.350 M HNO₂ in order to produce a buffer with a pH of 3.50?

$$\text{pH} = 3.50; [\text{H}_3\text{O}^+] = 3.2 \times 10^{-4} \text{ M}$$

$$K_{\text{a}} = \frac{[\text{NO}_2^-][\text{H}_3\text{O}^+]}{[\text{HNO}_2]} = 5.13 \times 10^{-4} = \frac{[\text{NO}_2^-] 3.2 \times 10^{-4}}{0.350}; [\text{NO}_2^-] = 0.56 \text{ M}$$

$$\frac{0.56 \text{ mol NaNO}_2}{\text{L}} \times \frac{68.996 \text{ g NaNO}_2}{1 \text{ mol NaNO}_2} \times \frac{0.2500 \text{ L}}{1} = 9.7 \text{ g NaNO}_2$$

19. Describe the preparation of 250.0 mL of a buffer with a pH of 5.000 from 0.1500 M $\text{HC}_2\text{H}_3\text{O}_2$ and 0.2000 M $\text{NaC}_2\text{H}_3\text{O}_2$.

$$K_a = \frac{[\text{C}_2\text{H}_3\text{O}_2^-][\text{H}_3\text{O}^+]}{[\text{HC}_2\text{H}_3\text{O}_2]} = 1.74 \times 10^{-5} = \frac{[\text{C}_2\text{H}_3\text{O}_2^-]1.00 \times 10^{-5}}{[\text{HC}_2\text{H}_3\text{O}_2]}; \frac{[\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} = 1.74$$

Let x = mL of $\text{HC}_2\text{H}_3\text{O}_2$ and y = mL of $\text{NaC}_2\text{H}_3\text{O}_2$; $x + y = 250.0$; $x = 250.0 - y$

$$[\text{HC}_2\text{H}_3\text{O}_2] = \frac{x \cdot 0.1500}{250.0} = \frac{(250.0 - y) \cdot 0.1500}{250.0}; [\text{C}_2\text{H}_3\text{O}_2^-] = \frac{y \cdot 0.2000}{250.0}$$

$$\frac{\frac{y \cdot 0.2000}{250.0}}{\frac{(250.0 - y) \cdot 0.1500}{250.0}} = \frac{y \cdot 0.2000}{(250.0 - y) \cdot 0.1500} = 1.74$$

$$0.2000y = 65.2 - 0.261y; y = 142; x = 250.0 - 142 = 108$$

Mix 108 mL of 0.1500 M $\text{HC}_2\text{H}_3\text{O}_2$ with 142 mL of 0.2000 M $\text{NaC}_2\text{H}_3\text{O}_2$.

20. 25.00 mL of 0.1500 M HCl are titrated with 0.3000 M NaOH. Calculate the pH after 7.50 mL of NaOH are added.

$$[\text{H}_3\text{O}^+] = 0.1500 \text{ M} \times \frac{25.00 \text{ mL}}{32.50 \text{ mL}} = 0.1154 \text{ M}$$

$$[\text{OH}^-] = 0.3000 \text{ M} \times \frac{7.50 \text{ mL}}{32.50 \text{ mL}} = 0.0692 \text{ M}$$

$$[\text{H}_3\text{O}^+] = 0.1154 \text{ M} - 0.0692 \text{ M} = 0.0462 \text{ M}; \text{pH} = 1.34$$

21. 50.00 mL of 0.2000 M $\text{HC}_2\text{H}_3\text{O}_2$ are titrated with 0.4000 M NaOH. Calculate the pH after 30.00 mL of NaOH are added.

$$[\text{HC}_2\text{H}_3\text{O}_2] = 0.2000 \text{ M} \times \frac{50.00 \text{ mL}}{80.00 \text{ mL}} = 0.1250 \text{ M}$$

$$[\text{OH}^-] = 0.4000 \text{ M} \times \frac{30.00 \text{ mL}}{80.00 \text{ mL}} = 0.1500 \text{ M}$$

$$[\text{OH}^-] = 0.1500 \text{ M} - 0.1250 \text{ M} = 0.0250 \text{ M}; \text{pOH} = 1.60; \text{pH} = 12.40$$

22. A 125.0 mL sample is removed from a water solution saturated with Ca(OH)_2 at 25.0°C . The water is completely evaporated from the solution and a deposit of 0.1030 g of Ca(OH)_2 is obtained. Calculate the K_{sp} of Ca(OH)_2 at 25.0°C .

$$\frac{0.1030 \text{ g Ca(OH)}_2}{1} \times \frac{1 \text{ mol Ca(OH)}_2}{74.093 \text{ g Ca(OH)}_2} = 0.001390 \text{ mol Ca(OH)}_2$$

$$\frac{0.001390 \text{ mol Ca(OH)}_2}{1} \times \frac{1 \text{ mol Ca}^{2+}}{1 \text{ mol Ca(OH)}_2} \times \frac{1}{0.1250 \text{ L}} = 0.01112 \text{ M Ca}^{2+}$$

$$\frac{0.001390 \text{ mol Ca(OH)}_2}{1} \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ca(OH)}_2} \times \frac{1}{0.1250 \text{ L}} = 0.02224 \text{ M OH}^-$$

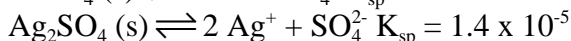
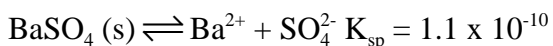
$$K_{\text{sp}} = [\text{Ca}^{2+}][\text{OH}^-]^2 = 0.01112 \times (0.02224)^2 = 5.502 \times 10^{-6}$$

23. It is desired to separate Ba^{2+} from Ag^+ by the slow addition of SO_4^{2-} to a solution in which the concentration of Ba^{2+} is 0.002500 M and the concentration of Ag^+ is 0.0001250 M.

(a) Which ion will precipitate first and at what $[\text{SO}_4^{2-}]$?

(b) What will be the concentration of the first ion to precipitate when the second ion just begins to precipitate?

(c) Is the separation effective?



$$(a) \quad K_{\text{sp}} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = 1.1 \times 10^{-10} = 0.002500 [\text{SO}_4^{2-}]; \quad [\text{SO}_4^{2-}] = 4.4 \times 10^{-8} \text{ M}$$

$$K_{\text{sp}} = [\text{Ag}^+]^2[\text{SO}_4^{2-}] = 1.4 \times 10^{-5} = (0.0001250)^2 [\text{SO}_4^{2-}]; \quad [\text{SO}_4^{2-}] = 896 \text{ M}$$

Ba^{2+} will precipitate first when $[\text{SO}_4^{2-}] = 4.4 \times 10^{-8} \text{ M}$

$$(b) \quad [\text{Ba}^{2+}] 896 = 1.1 \times 10^{-10}; \quad [\text{Ba}^{2+}] = 1.2 \times 10^{-13} \text{ M}$$

(c) Yes.

24. Eight drops of 0.01500 M NaCl are added to 175.0 mL of 0.007500 M AgNO₃. Will AgCl precipitate? 20 drops equal 1 mL.

$$\frac{8 \text{ d}}{1} \times \frac{1 \text{ mL}}{20 \text{ d}} = 0.40 \text{ mL}$$

$$\frac{0.01500 \text{ mol NaCl}}{\text{L}} \times \frac{0.40 \text{ mL}}{175.4 \text{ mL}} \times \frac{1 \text{ mol Cl}^-}{1 \text{ mol}} = 3.4 \times 10^{-5} \text{ M Cl}^-$$

$$\frac{0.007500 \text{ mol AgNO}_3}{\text{L}} \times \frac{175.0 \text{ mL}}{175.4 \text{ mL}} \times \frac{1 \text{ mol Ag}^+}{1 \text{ mol AgNO}_3} = 0.007483 \text{ M Ag}^+$$

$$Q = [\text{Ag}^+][\text{Cl}^-] = 3.4 \times 10^{-5} \times 0.007483 = 2.6 \times 10^{-7}$$

$Q > K_{\text{sp}}$ therefore AgCl (s) will form

25. What concentration of C₂H₅NH₃⁺ must be maintained in order to prevent the precipitation of Mg(OH)₂ in a solution that is 0.01250 M in Mg(NO₃)₂ and 0.003500 M in C₂H₅NH₂?

$$K_{\text{sp}} = [\text{Mg}^{2+}][\text{OH}^-]^2 = 1.8 \times 10^{-11} = 0.01250[\text{OH}^-]^2; [\text{OH}^-] = 3.8 \times 10^{-5} \text{ M}$$

$$K_{\text{b}} = \frac{[\text{C}_2\text{H}_5\text{NH}_3^+][\text{OH}^-]}{[\text{C}_2\text{H}_5\text{NH}_2]} = 4.4 \times 10^{-4} = \frac{[\text{C}_2\text{H}_5\text{NH}_3^+]3.8 \times 10^{-5}}{0.003500}; [\text{C}_2\text{H}_5\text{NH}_3^+] = 0.041 \text{ M}$$

26. What is the minimum concentration of NH₃ required to prevent AgCl from precipitating in a solution that is 0.0002500 M in AgNO₃ and 0.00009500 M in NaCl?

$$K_{\text{sp}} = [\text{Ag}^+][\text{Cl}^-] = 1.8 \times 10^{-10} = [\text{Ag}^+]0.00009500; [\text{Ag}^+] = 1.9 \times 10^{-6} \text{ M}$$

$$K_{\text{f}} = \frac{[\text{Ag}(\text{NH}_3)_2^+]}{[\text{Ag}^+][\text{NH}_3]^2} = 1.6 \times 10^7 = \frac{0.0002500}{1.9 \times 10^{-6} [\text{NH}_3]^2}$$

$$[\text{NH}_3] = 2.9 \times 10^{-3} \text{ M}$$

$$\frac{0.0002500 \text{ mol Ag}^+}{\text{L}} \times \frac{2 \text{ mol NH}_3}{1 \text{ mol Ag}^+} = 0.0005000 \text{ M NH}_3$$

$$[\text{NH}_3] = 0.0029 \text{ M} + 0.0005000 \text{ M} = 0.0034 \text{ M}$$