General Chemistry II Jasperse
Buffers/Titrations/Solubility. Extra Practice Problems

General Types/Groups of problems:

Buf

Buffers General	p1		Titration Graphs and Recognition	p10
What Kind of Solution/pH at End?	p2		Titration Calculations	p11
Preparation and Recognition of Buffers	p4		pH Estimations/Calculations after acid/base are added (including at Equivalence Point)	p12
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Sol

Key Equations Given for Test:

For weak acids alone in water: $[H^+] = \sqrt{K_a x[WA]}$	For weak bases alone in water: $[OH^-] = \sqrt{K_b x[WB]}$
pZ= -logZ General definition for p of anything	pH + pOH = 14
$[H^+][HO^-] = 1.00 \times 10^{-14}$	$K_aK_b = 1.00 \times 10^{-14}$ for conjugate acid/base pair
For Buffer: $pH = pK_a + \log[base]/[acid]$ Henderson-Hasselbalch Equation	ΔS° = S° (products) – S° (reactants)
ΔG° = G° (products) – G° (reactants)	ΔG° = ΔH° –TΔS° (T in Kelvin)

BUFFERS

1. A solution that contains a weak acid and its conjugate base in roughly equal concentrations is __________

a. neither acidic or basic.
b. a half-acid solution.
[✓ c. a buffer.]
d. a heterogeneous mixture.
e. neutral.

2. Explain how a buffer solution manages to stabilize the pH against the addition of acid, base, or additional solvent (dilution).

Answer: A buffer consists of a weak acid and its conjugate base in roughly equal amounts. If acid is added to the solution, it is consumed by the conjugate base. If base is added to the solution, it is consumed by the weak acid. If the amounts are such that the ratio of conjugate base/weak acid concentrations doesn't change much, then the pH doesn't change much. Dilution does not affect the pH because this concentration ratio doesn't change upon dilution.

3. Research with biochemical systems commonly requires buffers because __________

a. that's just the way it is.
[✓ b. proteins have a critical pH dependence in their structure and function.]
c. proteins decompose into constituent amino acids outside a certain pH range.
d. proteins are buffers.
e. salts are involved.

4. What reaction occurs as a hydrochloric acid solution is added to a solution containing equal concentrations of acetic acid and sodium acetate?

buffer

a. $CH_3COOH + H^+ \rightarrow CH_3COOH_2^+$
[✓ b. $CH_3COO^- + H^+ \rightarrow CH_3COOH$]
c. $CH_3COOH + HCl \rightarrow CH_3COO^- + H_2Cl^+$
d. $2CH_3COO^- + 2H^+ \rightarrow CH_3COO + H_2$
e. $CH_3COOH + H^+ \rightarrow CH_3CO^+ + H_2O$

What kind of Solution/pH Do you Get at the End? How does Solution Change?

S
S
T
S

5. When the following chemicals are mixed, each in 1 liter of water, which would give an acidic pH at the end?

a) 1 mole of KOH and 1 mole of NaF
b) 1 mole of NH_4^+ and 1 mole of NaOH
[✓ c)] 1 mole of HCl and 1 mole of NH_3
d) 1 mole of KOH and 0.5 mole of HCl

$SB + WB$

$WA + OH^{\ominus} \rightarrow \boxed{WB}$

$HO^{\ominus} = \sqrt{K_b[\text{base}]}$

$SA + WB \rightarrow \boxed{WA}$

$H^+ = \sqrt{K_a[HA]}$

0.5 KOH

6. When the following chemicals are mixed, each in 1 liter of water, which would give a basic pH at the end?

a) 1 mole of KOH and 1 mole of HF → WB
b) 1.0 mole of KOH and 1.0 mole of HCl → neutral
c) 1 mole of HCl and 1 mole of NH_3 → WA
d) 0.5 mole of KOH and 1.0 mole of HCl → SA

S
S
T
S

7. When the following chemicals are mixed, each in 1 liter of water, which would give a basic pH at the end?

a) 1 mole of KOH and 1 mole of HF → WB
b) 1.0 mole of KOH and 1.0 mole of HCl
c) 1 mole of HCl and 1 mole of NH_3
d) 0.5 mole of KOH and 1.0 mole of HCl

WB buffer WA

8. Consider a solution initially containing 0.40 mol fluoride anion and 0.30 mol of hydrogen fluoride (HF). If 0.20 mol of HCl are added to this solution, which of the following statements is **FALSE**?

SA

a) You will still have a buffer solution at the end, since you'll still have both weak base and conjugate weak acid
b) The pH will have shifted to a lower pH
c) You'll have more moles of HF at the end than what you began with
[✓ d)] You will no longer have a buffer solution, since all of the weak base will have reacted with the HCl.
e) none of the above

	HCl +	$F^{\ominus}$ →	HF
I	.2	.4	.3
C	−.2	−.2	+.2
E	—	.2	.5

WB + WA buffer

S
S
T
S

9. Consider a solution initially containing 0.40 mol fluoride anion and 0.30 mol of hydrogen fluoride (HF). If 0.40 mol of HCl are added to this solution, which of the following statements is **FALSE**?

[✓ a)] You will still have a buffer solution at the end, since you'll still have significant amounts of both weak base and conjugate weak acid
b) The pH will have shifted to a lower pH
c) You'll essentially have a weak acid solution situation, with 0.7 mol HF at the end.
d) You will no longer have a buffer solution, since all of the weak base will have reacted with the HCl. The buffer capacity was exhausted.
e) none of the above

HCl +	$F^{\ominus}$ →	HF
.4	.4	.3
−.4	−.4	+.4
—	—	.7

HF

WA

$\ominus OH + HF \rightarrow F^{\ominus}$ WB + SB →

.4 .3 .4

G .1 — .7

10. Consider a solution initially containing 0.40 mol fluoride anion and 0.30 mol of hydrogen fluoride (HF). If 0.40 mol of NaOH are added to this solution, and the final volume is 1L, which of the following statements is **FALSE**?

T

a) You'll essentially have a strong base solution at the end, with 0.7 mol NaF but with 0.1 mol NaOH at the end. The moles of strong base will essentially dictate the pH.
b) The pH will have shifted to a higher, more basic pH
c) The final pH will be 13.
d) The final pH will be dominated by the NaF, so I'd need the K_b or pK_b in order to solve for the pH.
e) none of the above

Buffer

11. Consider a solution that contains 0.50 moles of KF and 0.50 moles of HF in 1.0 L of water. If 0.10 mol of NaOH is added to this buffer solution, the pH of the solution will get slightly ____________________. The pH does not change more drastically because the NaOH reacts with the ______________________ present in the buffer solution.

a) higher, KF
[✓ b) higher, HF]
c) lower, KF
d) lower, HF

T

12. Consider a solution that contains 0.50 moles of $NaNO_2$ and 0.50 moles of HNO_2 in 1.0 L of water. If 0.10 mol of HCl is added to this buffer solution, the pH of the solution will get slightly ____________________. The pH does not change more drastically because the ~~NaOH~~ HCl reacts with the ______________________ present in the buffer solution.

a) higher, $NaNO_2$
b) higher, HNO_2
[✓ c) lower, $NaNO_2$]
d) lower, HNO_2

13. Identify all the *correct* statements about an acid–base buffer solution.

I. It can be prepared by combining a strong acid with a salt of its conjugate base. F
II. It can be prepared by combining a weak acid with a salt of its conjugate base. T
III. It can be prepared by combining a weak base with its conjugate acid. T
?: IV. The pH of a buffer solution does not change when the solution is diluted. T
V. A buffer solution resists changes in its pH when an acid or base is added to it. T

T

a. I, II, and IV
b. II, III, and V
[✓ c. II, III, IV, and V]
d. I, II, IV, and V
e. II, III, and IV

$$pH = pKa + \log \boxed{\frac{base}{acid}}$$

Preparation and Recognition of Buffer Systems

14. Which of the following is *not* a buffer system? A solution containing roughly equal concentrations of __________

- a. fluoride ion and hydrofluoric acid. WB + WA
- [✓ b.] bromide ion and hydrobromic acid. non-base + Strong acid
- c. phosphate ion and hydrogen phosphate ion. WB + WA
- d. carbonate ion and hydrogen carbonate ion. WB + WA
- e. phosphoric acid and dihydrogen phosphate ion. WA + WB

$1\ SA + 2\ WB \rightarrow WA$

$\boxed{1\ WB \quad 1\ WA}$

15. Which of the following *can* be mixed together in water to produce a buffer solution?

- a. $HClO_4$ and $NaClO_4$ SA + neutral
- b. HNO_3 and $NaNO_3$ SA + neutral
- c. H_2SO_4 and $NaHSO_4$ SA + WA
- [✓ d.] $\overset{WA}{H_3PO_4}$ and $\overset{WB}{NaH_2PO_4}$ WA + WB
- e. HCl and NaCl SA + neutral

16. Which one of the following would make the best buffer? (Ac = acetate, CH_3CO_2)

- a. a solution of hydrochloric acid and sodium chloride, HCl and NaCl SA + neutral
- b. a solution of acetic acid and ammonia, NaAc and NH_3 WB + WB
- c. a solution of acetic acid and ammonium chloride, HAc and NH_4Cl WA + WA
- d. a solution of sodium acetate and ammonium chloride, NaAc and NH_4Cl WB + WA but NOT conjugates
- [✓ e.] a solution of ammonia and ammonium chloride, NH_3 and NH_4Cl WB + WB and Conjugates

17. Which combination of solutions is the best choice for making a buffer solution?

- WB + WA a. equal volumes of 1 *M* ammonia (NH_3) and 0.001 *M* ammonium chloride (NH_4Cl) ← little
- SA + SB b. equal volumes of 0.5 *M* hydrochloric acid (HCl) and 0.5 *M* sodium hydroxide (NaOH)
- SA + neut c. equal volumes of 0.5 *M* hydrochloric acid (HCl) and 0.5 *M* sodium chloride (NaCl)
- WA + WB [✓ d.] equal volumes of 2 *M* ammonia (NH_3) and 1 *M* hydrochloric acid (HCl)
- WA + SA e. equal volumes of 2 *M* ammonium chloride (NH_4Cl) and 1 *M* hydrochloric acid (HCl)

$$\begin{array}{ccccc} NH_3 & + & HCl & \rightarrow & NH_4^+ \\ 2 & & 1 & & \\ -1 & & -1 & & +1 \\ \hline 1 & & - & & 1 \\ WB & & & & WA \end{array}$$

18. Which of the following would be the best choice for preparing a buffer with a pH = 8.0?

- a. a solution of formic acid and sodium formate, $K_a = 1.8 \times 10^{-4}$
- b. a solution of acetic acid and sodium acetate, $K_a = 1.8 \times 10^{-5}$
- [✓ c.] a solution of hypochlorous acid and sodium hypochlorite, $K_a = 3.5 \times 10^{-8}$
- d. a solution of boric acid and sodium borate, $K_a = 5.8 \times 10^{-10}$
- e. All of these solutions would be equally good choices for making this buffer.

$$\boxed{pH = pK_a} + \log\frac{base}{acid}$$

19. When placed in 1 L of water, which of the following combinations would give a buffer solution? (Remember, in some cases they might react with each other…)

- [✓ 1)] 0.5 mol HClO and 0.5 mol NaClO
- ~~2) 0.5 mol HBr and 0.5 mol NaF~~ → HF
- [✓ 3)] 0.5 mol HBr and 1.0 mol NaF → 0.5 HF + 0.5 NaF
- ~~4) 0.5 mol HBr and 1.0 mol NaOH~~ OH

a) 1 only
b) 1 and 2 only
[✓ c)] 1 and 3 only
d) 3 and 4 only
e) all would give buffer solutions

[redacted] Calculations

S
S
T
S

Finding pH

20. Calculate the pH of a solution that is 0.30 *M* in ammonia (NH_3) and 0.20 *M* in ammonium chloride (NH_4Cl, $K_a = 5.62 \times 10^{-10}$).

WB

WA

Answer: 9.43

$$\boxed{pH = pK_a + \log \frac{base}{acid}}$$

$$= 9.25 + \log\left(\frac{.3}{.2}\right) = \boxed{9.43}$$

Steps:
1. Recognize: WA + WB = buffer
2. Solve: HH equation
3. Need: pKA
4. Check: Does it make sense?

-pH is close to pKa, basic side

21. Calculate the pH of a solution containing 0.40 mol fluoride anion and 0.30 mol of hydrogen fluoride (HF). (HF, $K_a = 7.2 \times 10^{-4}$)? → $pK_a = 3.14$

a) 0.20 mol
b) 0.30 mol
c) 0.40 mol
d) 0.50 mol
e) none of the above

$$pH = 3.14 + \log \frac{.4}{.3}$$

$$\boxed{pH = 3.27}$$

Steps:
1. Recognize: WA + WB = buffer
2. Solve: HH equation
3. Need: pKA
4. Check: Does it make sense?

-pH is close to pKa, basic side

S
S
T
S

22. Calculate the pH of a solution that's 0.65M in NaF and 0.75M in HF. (HF, $K_a = 7.2 \times 10^{-4}$)?

ANSWER: 3.08

Steps:
1. Recognize: WA + WB = buffer
2. Solve: HH equation
3. Need: pKA
4. Check: Does it make sense?

-pH is close to pKa, acid side

$pK_a = 3.14$

$$pH = 3.14 + \log\left(\frac{.65}{.75}\right)$$

$$= 3.08$$

23. Calculate the pH of a solution that's 0.65M in $NaNO_2$ and 0.40M in HNO_2. ($NaNO_2$, $K_a = 4.0 \times 10^{-4}$)? 3.40 pK_a

ANSWER: 3.61

$$pH = 3.40 + \log\left(\frac{.65}{.40}\right) = 3.609$$

Steps:
1. Recognize: WA + WB = buffer
2. Solve: HH equation
3. Need: pKA
4. Check: Does it make sense?

Finding Amount of Base or Acid or Ratio

24. To simulate the pH of blood, which is 7.4, an undergraduate researcher in a biology lab produced a buffer solution by dissolving sodium dihydrogen phosphate (NaH_2PO_4, $K_a = 6.2 \times 10^{-8}$) and sodium hydrogen phosphate (Na_2HPO_4) together in an aqueous solution. What mole ratio of Na_2HPO_4/NaH_2PO_4 did she need to use?

WA

7.21

WB

a. 1.2
b. 1.6
c. 0.90
d. 1.0
e. 0.96

S
S
T
S

1. Recognize: => HH equation
2. Need: pKa, then Base/Acid
3. Check: Does it make sense?

Basic, so Base/Acid => >1

$$7.4 = 7.21 + \log \frac{base}{acid}$$

$$\boxed{\log x = 0.19}$$

$$10^{0.19} = 1.56$$

25. What is the pH of a buffer solution where [HA] = [A⁻]?

a. pH = 1
b. pH = K_a
c. pH = pK_a
d. pH = pOH
e. pH = 7.0

log → 0

$$\boxed{pH = pK_a} + \cancel{\log \frac{b}{a}}$$

T

26. A buffer system is set up with [HA] = 2[A$^-$]. If pK_a = 5.5, what is the pH of the buffer?

[✓ a. 5.2]
b. 5.8
c. 7.5
d. 3.5
e. 7.0

$pH = \cancel{0}5.5 + \log \cancel{2}\left(\frac{1}{2}\right)$

Finding pKa
↓
Ka

27. The pK_a of a weak acid was determined by measuring the pH of a solution containing the weak acid at 0.30 M and its conjugate base at 0.30 M. The measured pH was 8.0. What is the pK_a of the weak acid?

[✓ a. 8.0]
b. 7.8
c. 7.6
d. 7.0
e. 7.4

$8 = pK_a + \cancel{\log\left(\frac{.30}{.30}\right)}$

28. The pK_a of a weak acid was determined by measuring the pH of a solution containing the weak acid at 0.40 M and its conjugate base at 0.60 M. The measured pH was 7.8. What is the pK_a of the weak acid?

a. 8.0
b. 7.8
[✓ c. 7.6]
d. 7.0
e. 7.4

$7.8 = x + \log\left(\frac{.60}{.40}\right)$

T

29. The pK_a of a weak acid was determined by measuring the pH of a solution containing the weak acid at 0.40 M and its conjugate base at 0.20 M. The measured pH was 9.8. What is the pK_a of the weak acid?

[✓ a. 10.1]
b. 9.8
c. 9.5
d. 10.4
e. None of the above

$9.8 = x + \log\left(\frac{.2}{.4}\right)$

$9.8 = x - .30$

Finding Amount of Base or Acid or Ratio

30. How many moles of sodium acetate must be added to 500 mL of 0.25 M acetic acid solution to produce a buffer with a pH of 4.94? The pK_a of acetic acid is 4.74.

a. 0.011 moles
b. 0.021 moles
c. 0.125 moles
[✓ d. 0.198 moles]
e. 0.206 moles

$(.500\,L)\left(\frac{.25\,mol}{1\,L}\right) = .125\,mol$

$4.94 = 4.74 + \log\left(\frac{b}{a}\right)$

$0.2 = \log x$

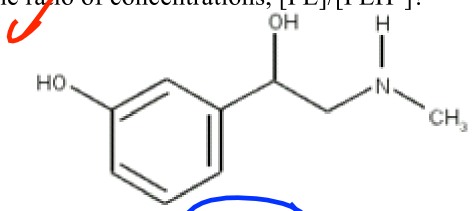
$x = \boxed{1.58 = \frac{base}{acid}} = \frac{x}{.125}$

$\boxed{x = .198\,mol}$

T

31. Phenylephrine (PE, see the structure below) is a nasal decongestant and is the active ingredient in Sudafed, which contains phenylephrine hydrochloride (PEHCl). This conjugate acid of phenylephrine (PEH$^+$) has a pK_a = 5.5. At a physiological pH of 7.4, what is the ratio of concentrations, [PE]/[PEH$^+$]?

WA



a. 6.7
b. ~~0.01~~
c. ~~0.14~~
[✓ d. 79]
e. 21

$pH = pK_a + \log \frac{b}{a}$

$\boxed{7.4 = 5.5} + \log x$

$1.9 = \log x$

$\boxed{x = 79}$

$x = 10^{1.9}$

Disrupted Buffers. What Happens after Acid or Base is added.

32. Consider a solution initially containing 0.40 mol fluoride anion and 0.30 mol of hydrogen fluoride (HF). How many moles of hydrogen fluoride are present after addition of 0.20 mol of HCl to this solution?

T

Note: Moles Given Directly

Note: What reacts with what
What runs out first?
What's left at the end?

a) 0.20 mol
b) 0.30 mol
c) 0.40 mol
d) 0.50 mol
e) none of the above

$$\begin{array}{lcccc} HCl & + & F^{\ominus} & \longrightarrow & HF \\ .2 & & .4 & & .3 \\ -.2 & & -.2 & & +.2 \\ \hline & & .2 \text{ WB} & & .5 \text{ WA} \end{array}$$

33. Consider a solution initially containing 0.50 mol ammonia (NH$_3$) and 0.30 mol of ammonium ion (NH$_4^+$). How many moles of ammonia and how many moles of ammonium ion are present after addition of 0.20 mol of HCl to this solution?

Text

a) 0.30 mol NH$_3$, 0.50 mol NH$_4^+$
a) 0.50 mol NH$_3$, 0.50 mol NH$_4^+$
a) 0.30 mol NH$_3$, 0.20 mol NH$_4^+$
a) 0.70 mol NH$_3$, 0.10 mol NH$_4^+$
e) none of the above

$$\begin{array}{lcccc} HCl & + & NH_3 & \longrightarrow & {}^{\oplus}NH_4 \\ .2 & & .5 & & .3 \\ -.2 & & -.2 & & +.2 \\ \hline & & .3 \text{ WB} & & .5 \text{ WA} \end{array}$$

T

34. Consider a solution initially containing 0.400 mol fluoride anion and 0.300 mol of hydrogen fluoride (HF). How many moles of hydrogen fluoride are present after addition of 70.0 mL of 0.600*M* HCl to this solution?

a) 0.400 mol fluoride, 0.300 mol HF
b) 0.442 mol fluoride, 0.258 mol HF
c) 0.358 mol fluoride, 0.342 mol HF
d) 0.213 mol fluoride, 0.567 mol HF
e) none of the above

$$\begin{array}{lcccc} HCl & + & F^{\ominus} & \longrightarrow & HF \\ .042 & & .4 \text{ mol} & & .3 \text{ mol} \\ -.042 & & -.042 & & +.042 \\ \hline & & .358 & & .342 \end{array}$$

$$\text{mol HCl} = (.070\,\cancel{L})\left(\frac{.6\text{ mol}}{1\,\cancel{L}}\right) = .042\text{ mol}$$

Note: Moles Given Indirectly

35. Consider a solution initially containing 0.500 mol ammonia (NH$_3$) and 0.300 mol of ammonium ion (NH$_4^+$). How many moles of ammonia and how many moles of ammonium ion are present after addition of 40 mL of 0.800*M* NaOH to this solution?

a) 0.532 mol NH$_3$, 0.268 mol NH$_4^+$
a) 0.532 mol NH$_3$, 0.332 mol NH$_4^+$
a) 0.468mol NH$_3$, 0.268 mol NH$_4^+$
a) 0.700 mol NH$_3$, 0.100 mol NH$_4^+$
e) none of the above

$$\begin{array}{lcccc} {}^{\ominus}OH & + & {}^{\oplus}NH_4 & \longrightarrow & NH_3 \\ .032 & & .300 & & .500 \text{ mol} \\ -.032 & & -.032 & & +.032 \\ \hline & & .268 & & .532 \end{array}$$

$$\text{mol } {}^{\ominus}OH = (.040\,\cancel{L})\left(\frac{.800\text{ mol}}{1\,\cancel{L}}\right) = .032\text{ mol}$$

T

36. Consider a solution initially containing 0.40 mol fluoride anion and 0.30 mol of hydrogen fluoride (HF). **What is the pH** after addition of 0.20 mol of HCl to this solution? (HF, $K_a = 7.2 \times 10^{-4}$)? $\rightarrow pK_a = 3.14$

ANSWER: 2.74

$$\begin{array}{lcccc} HCl & + & F^{\ominus} & \longrightarrow & HF \\ .2 & & .4 & & .3 \\ -.2 & & -.2 & & +.2 \\ \hline & & .2\text{ mol} & & .5\text{ mol} \end{array}$$

buffer ⇒ HH

$$pH = 3.14 + \log\left(\frac{.2}{.5}\right)$$

$$\boxed{pH = 2.74}$$

Text

Note: In the problems on this page, there was never enough SA or SB added to exhaust the buffer. You STILL had a buffer at the end. But, there can be other situations in which the buffer capacity is more than exhausted. In which case you will NOT end with a buffer, and you won't end up solving by Henderson-Hasselbalch.

37. Consider a solution initially containing 0.400 mol fluoride anion and 0.300 mol of hydrogen fluoride (HF). **What is the pH** after addition of 70.0 mL of 0.600*M* HCl to this solution? (HF, $K_a = 7.2 \times 10^{-4}$)? $\rightarrow pK_a = 3.14$

T

ANSWER: 3.16

$HCl + F^{\ominus} \rightarrow HF$

.042 .400 mol .300 mol

.358 .342 mol

$pH = 3.14 + \log\left(\frac{.358}{.342}\right)$

$\boxed{= 3.16}$

38. Consider a solution initially containing 0.50 mol ammonia (NH_3) and 0.30 mol of ammonium ion (NH_4^+). **What is the pH** after addition of 0.20 mol of HCl to this solution? (NH_4^+, $K_a = 5.6 \times 10^{-10}$)? $\rightarrow pK_a = 9.25$

ANSWER: 9.03

$HCl + NH_3 \rightarrow \overset{\oplus}{N}H_4$

.2 .5 .3

.3 .5

$pH = 9.25 + \log\left(\frac{.3}{.5}\right)$

$\boxed{= 9.03}$

39. Consider a solution initially ~~containing 0.500 mol~~ ammonia (NH_3) and 0.300 mol of ammonium ion (NH_4^+). **What is the pH** after addition of 40 mL of 0.800*M* NaOH to this solution? (NH_4^+, $K_a = 5.6 \times 10^{-10}$)? $pK_a = 9.25$

T

ANSWER: 9.55

.032 mol

$^{\ominus}OH + \overset{\oplus}{N}H_4 \rightarrow NH_3$

.032 .3 .5

−.032 +.032

.268 .532

40. Calculate the pH of a solution originally containing 0.20 mol of cyanic acid HCNO (WA) following addition of 80 mL of 1.00 M NaOH (SB). (K_a of HCNO = 3.5 x 10^{-4}). The initial volume of the cyanic acid solution was 920 mL, so the final combined volume at the end is 1.0 L.

[✓ a) 3.28]
b) 3.39
c) 3.46
d) 3.64
e) none of the above

$pK_a = 3.46$

$HO^{\ominus} + \underset{WA}{HCNO} \rightarrow \underset{WB}{^{\ominus}CNO}$

I .080 .200 ~0

E −.080 −.080 +.080

.120 WA .080 WB

$pH = 3.46 + \log\frac{.080}{.120}$

⇒ buffer

$mol\ ^{\ominus}OH = \frac{.080\ L}{} \left| \frac{1\ mol}{1\ L} \right. = .080\ mol$

41. Consider a solution initially containing 0.300 mol of hydrogen fluoride (HF). **How many grams of NaF (42.0 g/mol) would be needed** to set the pH = 3.00? (HF, $K_a = 7.2 \times 10^{-4}$)? $\rightarrow 3.14\ pK_a$

Making a Buffer.

T

ANSWER: 9.07g

$\boxed{9.1\,g}$

① pH + pKa (Ka) → b/a ratio
② Given moles acid ⇒ moles base
↓ mw
g base
③

$3.00 = 3.14 + \log\left(\frac{b}{a}\right)$

$-.14 = \log x$

$x = \left(\frac{.724\ mol\ base}{1\ mol\ \cancel{acid}}\right)\left(.300\ \cancel{mol\ acid}\right) = (.217\ mol\ base)\left(\frac{42\,g}{1\ mol}\right)$

$\boxed{= 9.1\,g}$

CO_3^{2-} PO_4^{3-} SO_4^{2-}

H_2CO_3 H_3PO_4 H_2SO_4

Di- or triprotic acids. A di- or triprotic base would be one that can absorb 2 or 3 H's.

Titration Related Problems

S S T S

42. In a titration of monoprotic acids and bases, there is a large change in pH __________
    - a. at the point where pH = pK_a of the acid.
    - b. when the volume of acid is exactly equal to the volume of base.
    - c. when the concentration of acid is exactly equal to the concentration of base.
    - [✓ d.] when the number of moles of acid is exactly equal to the number of moles of base.
    - e. at the point where pH = pK_b of the base.

43. At what point in the following titration curve for a weak acid being titrated with a strong base is the pH equal to the pK_a of the acid? The *x*-axis scale goes from 0.0 mL to 20.0 mL. The sharp rise is at 10.0 mL.

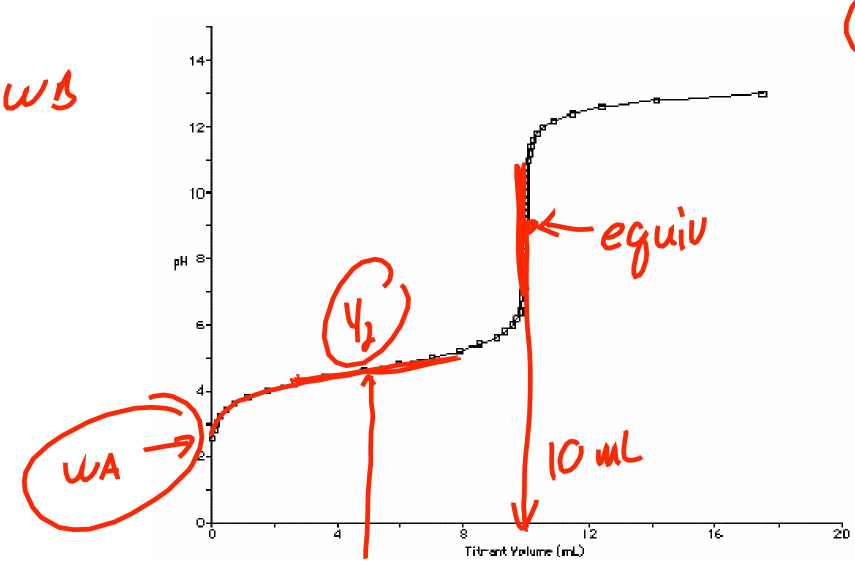
pH = pK + log (Base/acid)

Base/acid ratio = 1 at half-way-to-equivalence point

Note: should be able to diagnose the titration:
1. Starting acid or base?
2. Equivalence neutral, acidic, or basic?

WA → WB

S S T S



- a. 0.0 mL
- [✓ b.] 5.0 mL
- c. 9.0 mL
- d. 10.0 mL
- e. 18.0 mL

44. When an acetic acid solution (WA) is titrated with sodium hydroxide, the slope of the titration curve (pH vs volume of NaOH added) increases when sodium hydroxide is first added. This change shows that __________.
    - a. nothing is happening during this part of the titration.
    - b. the reaction is very slow during this part of the titration.
    - c. a more concentrated solution of NaOH needs to be present to initiate the reaction.
    - [✓ d.] acetic acid is being converted to sodium acetate.
    - e. the pH is not affected until all the acetic acid is consumed.

S S T S

45. Halfway to the equivalence point in a titration curve of a weak acid with a strong base, __________
    - a. nothing is happening yet.
    - [✓ b.] the pH = pK_a of the weak acid.
    - c. pH = 3.5 exactly.
    - d. pH = pK_a of the indicator.
    - e. the pH has not yet changed.

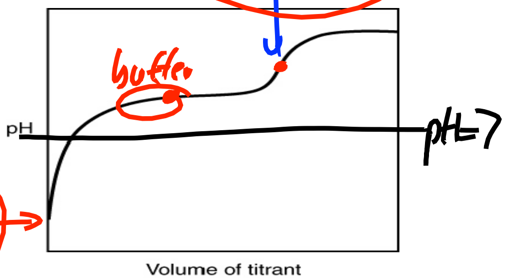
WB → WA

46. An initial pH of 9.5 and an equivalence point at pH 4.5 correspond to a titration curve for a

a) strong acid to which strong base is added
b) strong base to which strong acid is added
c) weak acid to which strong base is added
[✓ d)] weak base to which strong acid is added

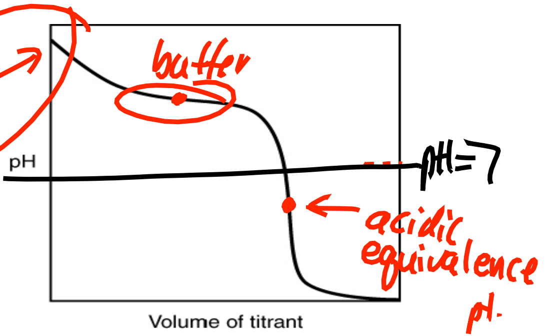
Titration Graphs and Recognition

47. The following titration curve is most likely to be associated with ___________



- a. the titration of a strong acid with a strong base titrant.
- [✓ b.] the titration of a weak acid with a strong base titrant.
- c. the titration of a strong base with a strong acid titrant.
- d. the titration of a weak base with a strong acid titrant.

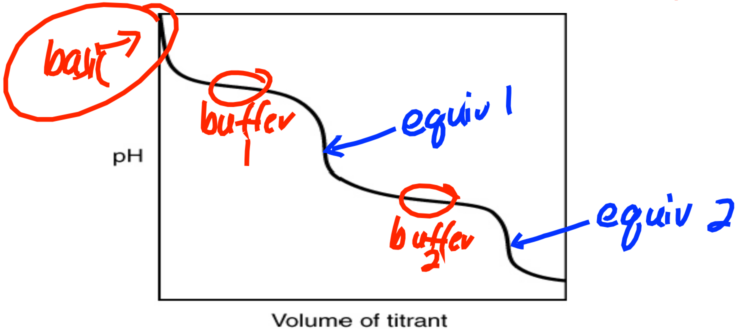
48. The following titration curve is most likely to be associated with



- a. the titration of a strong acid with a strong base titrant.
- b. the titration of a weak acid with a strong base titrant.
- c. the titration of a strong base with a strong acid titrant.
- [✓ d.] the titration of a weak base with a strong acid titrant.

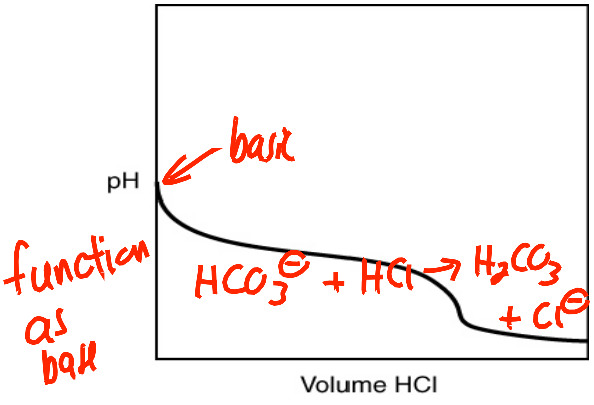
$CO_3^{2-} \rightarrow HCO_3^{\ominus} \rightarrow H_2CO_3$

49. What is indicated by the shape of the titration curve?

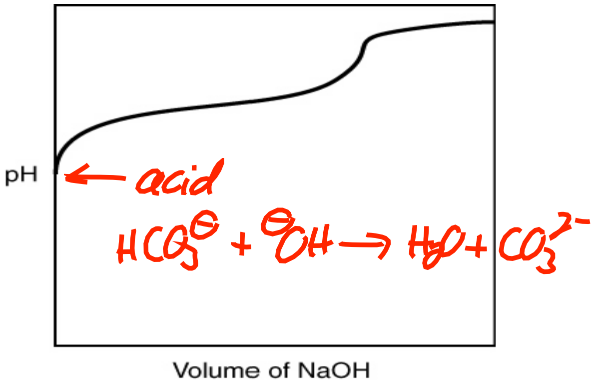


- a. A diprotic acid was titrated with a strong base.
- b. A triprotic acid was titrated with a strong base.
- [✓ c.] A diprotic base was titrated with a strong acid.
- d. A triprotic base was titrated with a strong acid.
- e. A strong acid was titrated with a strong base.

50. A 0.500 g sample of an unknown substance was titrated with a 0.1 *M* HCl solution. Another 0.500 g sample of it was titrated with a 0.1 *M* NaOH solution. The resulting titration curves are illustrated here. Given the following possibilities, what is the sample?



$HCO_3^{\ominus} + HCl \rightarrow H_2CO_3 + Cl^{\ominus}$



$HCO_3^{\ominus} + {}^{\ominus}OH \rightarrow H_2O + CO_3^{2-}$

- a. Na_2CO_3
- [✓ b.] $NaHCO_3$
- c. H_2CO_3
- d. CO_2
- e. There is no way to tell.

or $\boxed{V_a = V_b\left(\frac{M_b}{M_a}\right)}$

1. Equivalence point: Moles acid = Moles base
2. To know one is to know the other.
3. One is usually provided indirectly
moles = M x Volume in Liters, or from grams and mw)

mol = $\frac{M}{L}$ × L or mol = (g)($\frac{mol}{g}$)
Molarity volume g MW

$M_aV_a = M_bV_b$
so $\boxed{M_a = M_b\left(\frac{V_b}{V_a}\right)}$

Titration Calculations and Related Calculations.

S
S
T
S

51. A solution of hydrochloric acid (HCl, 25.00 mL) was titrated to the equivalence point with 34.55 mL of 0.1020 *M* sodium hydroxide. What was the concentration of the hydrochloric acid?

Approach 1: 1. Find moles base .003254 mol
2. Moles base = moles acid
3. Moles/volume (of the acid) = molarity
↳ .025 L

Approach 2: the Ratio method
Molarity of acid = Molarity of base x Volume Ratio

a. 0.07048 *M*
b. 0.1410 *M*
c. 0.2819 *M*
d. 0.0353 *M*
e. 0.0533 *M*

$\boxed{M_2 = M_1\left(\frac{V_1}{V_2}\right)}$

[HCl] = $(.1020)\left(\frac{34.55}{25.00}\right)$

$\frac{.003254 \text{ mol}}{.025 \text{ L}}$

Finding Unknown Molarity

52. A solution of hydrochloric acid (HCl, 40.00 mL) was titrated to the equivalence point with 22.0 mL of 0.320 *M* sodium hydroxide. What was the concentration of the hydrochloric acid?

$\frac{.00704 \text{ mol}}{.040}$

a. 0.07048 *M*
b. 0.178 *M*
c. 0.282 *M*
d. 0.0353 *M*
e. None of the above

$M = (.320 M)\left(\frac{22 \text{ mL}}{40 \text{ mL}}\right)$

53. A solution of sodium hydroxide (NaOH, 30.00 mL) was titrated to the equivalence point with 26.0 mL of 0.80 *M* HCl. What was the concentration of the sodium hydroxide solution?

a. 1.44 *M*
b. 0.178 *M*
c. 0.282 *M*
d. 0.693 *M*
e. None of the above

$[OH] = (.80)\left(\frac{26 \text{ mL}}{30 \text{ mL}}\right) =$

S
S
T
S

54. A solution of sulfuric acid (H_2SO_4, 25.00 mL) was titrated to the second equivalence point (both protons were removed) with 34.55 mL of 0.1020 *M* sodium hydroxide. What was the concentration of the sulfuric acid?

Polyprotic: Not 1:1 ratio

a. 0.07048 *M*
b. 0.1410 *M*
c. 0.2819 *M*
d. 0.0353 *M*
e. 0.0533 *M*

3.52×10^{-3} mol ^-OH

$\frac{1.76 \times 10^{-3} \text{ mol } H_2SO_4}{.025 \text{ L}}$

$1 H_2SO_4 + 2 NaOH \rightarrow 2H_2O + SO_4^{2-}$

55. What volume of 0.80 *M* NaOH will be required to titrate a 20.0 mL solution of 0.60M hydrochloric acid to the equivalence point?

Finding Unknown Volume

a. 15 mL
b. 20 mL
c. 25 mL
d. 30 mL
e. None of the above

$M_bV_b = M_aV_a$

$V_b = V_a\left(\frac{M_a}{M_b}\right)$

$= 20 \text{ mL}\left(\frac{.60}{.80}\right) = 15 \text{ mL}$

Using Grams instead of Volume to Provide Known Moles

56. What volume of 0.80 *M* HCl will be required to titrate 36.2 grams of NaBrO (fw = 118.9 g/mol) to the equivalence point?

mol b = $\frac{36.2 \text{ g}}{} \Big| \frac{1 \text{ mol}}{118.9 \text{ g}}$ = .304 mol

a. 150 mL
b. 381 mL
c. 258 mL
d. 308 mL
e. None of the above

x L = $(.304 \text{ mol})\left(\frac{1 \text{ L}}{.80 \text{ mol}}\right) =$

= .381 L

Harder One.
1. g => mol
2. 2:1 stoich
3. pH => [H+]
4. Mol => volume

need molarity

57. One brand of extra-strength antacid tablets contains 750 mg of calcium carbonate (100 g/mol) in each tablet. Stomach acid is essentially a hydrochloric acid solution. Is so much calcium carbonate really needed to neutralize stomach acid? Calculate the volume of stomach acid with a pH of 1.0 that one of these tablets could neutralize, and compare that value with the normal volume of stomach fluid, which usually is about 100 mL. One tablet can neutralize __________ mL of stomach acid at a pH of 1.0. (Remember, one carbonate can absorb not just one but two protons.)

$1CO_3^{2-} + 2H^+ \rightarrow H_2CO_3$

× mol CO_3^{2-} = $\frac{750 \text{ mg}}{} \Big| \frac{1 \text{ g}}{1000 \text{ mg}} \Big| \frac{1 \text{ mol}}{100 \text{ g}}$ = .0075 mol CO_3^{2-}

×2

.0150 mol HCl

a. 75
b. 150
c. 250
d. 15
e. 7.5

HCl: pH = 1.0 → $\frac{0.1 \text{ mol}}{\text{L}}$

× mL HCl = $\frac{.0150 \text{ mol}}{} \Big| \frac{1 \text{ L}}{0.1 \text{ mol}} \Big| \frac{1000 \text{ mL}}{1 \text{ L}}$ = $\boxed{150 \text{ mL}}$

pH Estimations or Calculations after acid or base are added (including at Equivalence Point)

S
S
T
S

58. Which of the following combinations would give a pH of 7.00 at the "equivalence point" (when equal moles of each have been added)? neutral

a) HCl + KF — HF + KCl acidic WA
b) HCN + NaOH — H2O + KCN basic WB
c) HF + HCl — No reaction, WA + SA = acidic
d) HCl + KOH — H2O + KCl neutral. Strong acid + strong base => neutral

59. Which of the following combinations would give a pH **above** 7.00 at the "equivalence point" (when equal moles of each have been added)? basic

a) HCl + KF — HF + KCl acidic
WA b) HCN + NaOH — H2O + NaCN basic WB
c) HF + HCl — No reaction, WA + SA = acidic
d) HCl + KOH — H2O + KCl neutral. Strong acid + strong base => neutral

60. Which of the following combinations would give a pH **below** 7.00 at the "equivalence point" (when equal moles of each have been added)? acid

a) HCl + KF WB — HF + KCl acidic
b) HCN + NaOH — H2O + NaCN basic
c) NH_3 + NaOH — No reaction, WB + SB = basic
d) HCl + KOH — H2O + KCl neutral. Strong acid + strong base => neutral

S
S
T
S

1. WA => WB at equivalence
2. For WB, need Kb and [WB]
3. To calculate [WB], need both moles and combined volume

$^{\ominus}OH \rightarrow pOH \rightarrow pH$

61. Glycolic acid, which is a monoprotic acid and a constituent in sugar cane, has a pK_a of 3.9. A 25.0 mL solution of glycolic acid is titrated to the equivalence point with 35.8 mL of 0.020 *M* sodium hydroxide solution. **What is the pH of the resulting solution at the equivalence point?**

WA

a. 5.10
b. 7.98
c. 8.72
d. 4.92
e. 9.08

$pK_a \rightarrow pK_b \rightarrow K_b$ 7.94×10^{-11}

WA → $\boxed{WB}$

60.8 mL = .0608 L

$[HO^-] = \sqrt{K_b \cdot [WB]} = \sqrt{(7.94\times10^{-11})(.0118)} = 9.67\times10^{-7} \rightarrow pOH = 6.01$

$\boxed{pH = 7.98}$

mol base $= \dfrac{7.16\times10^{-4}\text{ mol}}{.0608\text{ L}} = \boxed{.0118\text{ mol}}$

1. WB => WA at equivalence
2. For WA, need Ka and [WA]
3. To calculate [WA], need both moles and combined volume

62. Quinine is a weak base, with pK_b = 5.10. What is the pH if a 25.0 mL solution originally containing 0.125 moles of quinine is titrated with HCl to the equivalence point, and if the combined total volume at the end is 56.0 mL?

a. 5.10
b. 7.98
c. 8.72
d. 4.28
e. None of the above

WB → WA

$[H^+] = \sqrt{K_a \cdot [WA]}$

$pK_b = 5.10 \Rightarrow pK_a = 8.9 \Rightarrow \boxed{K_a = 1.26\times10^{-9}}$

81 mL

$\dfrac{.125\text{ mol}}{.081\text{ L}} = 1.54\text{ M}$

$H^+ = \sqrt{(1.26\times10^{-9})(1.54)} = 4.41\times10^{-5}$

$\boxed{pH = 4.35}$

1. Don't assume equivalence point unless told so. This one isn't!
2. Do "ICE" to figure out what you have.
3. Here, buffer, so use Henderson-Ha
4. Need pKa

63. A 25.0 mL solution of quinine was titrated with 1.00 *M* hydrochloric acid, HCl. It was found that the solution originally contained 0.125 moles of quinine. What was the pH of the solution after 50.00 mL of the HCl solution were added? Quinine is monobasic with pK_b = 5.10. → $pK_a = 8.90$

a. 5.10
b. 8.90
c. 8.72
d. 4.92
e. 9.08

mol HCl $= \left(\dfrac{1.0\text{ mol}}{1\text{ L}}\right).050\text{ L} = .05\text{ mol}$

$pH = 8.90 + \log\dfrac{.075}{.050} = 9.08$

	B	+ HCl →	$BH^{\oplus}$
I	.125	.050	~0
C	-.050	-.050	+.050
E	.075	—	.050

WB/WA ⇒ buffer

S
S
T
S

mol (80.0 mL) mol base (40.0 mL)

64. If you start with 80.0 mL of 0.40 M HNO_3, calculate the $[H^+]$ concentration following addition of 40.0 mL of 0.60 M KOH.

a) 0.0667 M
b) 1.00×10^{-7} M
c) 0.100 M
d) 1.50×10^{-13} M
e) none of the above

$HNO_3 + {}^{\ominus}OH \rightarrow H_2O + NO_3^{\ominus}$

$.032 \quad .024$

$-.024 \quad -.024$

$\frac{.008 \text{ mol}}{.120 L} = .0667$

Still had HNO_3 left

SA

1. What did I have at first? SA + SB
2. Find moles of each
3. What's left at end?
4. How do I solve?

(SA left this time, need moles and volume)

Need combined volume to find molarity

65. If you start with 80.0 mL of 0.40 M HNO_3, calculate the pH following addition of 40.0 mL of 0.80 M KOH.

a) 3.4
b) 7.0
c) 8.2
d) 11.6
e) none of the above

$HNO_3 + KOH \rightarrow KNO_3 + H_2O$

~~.032 .032~~

Equiv

1. What did I have at first? SA + SB
2. Find moles of each
3. What's left at end?
4. How do I solve?

(Nothing left this time = neutral)

S
T
S

66. If you start with 80.0 mL of 0.40 M HNO_3, calculate the pH following addition of 50.0 mL of 0.80 M KOH.

a) 3.4
b) 7.0
c) 12.8
d) 13.1
e) none of the above

$.032$

$-.032$

0

$.040$ mol

$-.032$

$\frac{.008 \text{ mol } {}^{\ominus}OH}{.130 L} = .0615$ M

pOH = -1.21

pH = 12.79

SB

1. What did I have at first? SA + SB
2. Find moles of each
3. What's left at end?
4. How do I solve?

(SB left this time, need moles and volume)

NEED COMBINED VOLUME TO CALCULATE Molarity

S
S
T
S

Solubility Problems

T

67. Write the reaction equation and the equilibrium constant expression that describes lead chloride ($PbCl_2$) dissolving in water.

Answer: $PbCl_2(s) \rightarrow 1Pb^{2+}(aq) + 2Cl^-(aq)$ $K_{sp} = [Pb^{2+}][Cl^-]^2$

x $2x$ $x \cdot 2x = x \cdot 4x^2 = 4x^3$

68. The solubility product for an insoluble salt with the formula MX_2 is written as __________, where x is the molar solubility.

a. $K_{sp} = x^2$
[✓ b. $K_{sp} = 4x^3$]
c. $K_{sp} = 4x^2$
d. $K_{sp} = 2x^3$
e. $K_{sp} = 2x^2$

$MX_2(s) \rightarrow 1M^{2+} + 2X^-$

x $2x$

$K = (x)(2x)^2 = 4x^3$

Find Ksp from Molar Solubility

T

69. When lead chloride ($PbCl_2$) is placed in otherwise pure water, enough dissolves such that the concentration of lead ions becomes 0.036 *M*. What is the K_{sp} for lead chloride ($PbCl_2$)?

a. $K_{sp} = 1.9 \times 10^{-6}$
b. $K_{sp} = 2.3 \times 10^{-5}$
c. $K_{sp} = 4.3 \times 10^{-6}$
[✓ d. $K_{sp} = 1.9 \times 10^{-4}$]
e. None of the above

$PbCl_2(s) \rightarrow 1Pb^{2+} + 2Cl^-$

$K = [Pb^{2+}][Cl^-]^2$

$= (.036)(.072)^2$

$\boxed{[Cl^-] = 2[Pb^{2+}]}$

70. Consider the following table of K_{sp} values.

Compound	K_{sp}
$PbCO_3$	7.4×10^{-14}
$ZnCO_3$	1.4×10^{-11}
$CePO_4$	6.3×10^{-16}
PbS	8.0×10^{-28}

Which one of the compounds shown in the table is the <u>least</u> soluble?

a) $PbCO_3$
b) $ZnCO_3$
c) $CePO_4$
d) PbS

Find Molar solubility from Ksp

71. The Ksp ~~solubility~~ of AgBr is 5.4×10^{-13} mol/L. What is the solubility ~~K_{sp}~~ of AgBr?

a) 5.4×10^{-13}
[✓ b) 7.3×10^{-7}]
c) 9.5×10^{-5}
d) 3.0×10^{-10}
e) none of the above

$AgBr(s) \rightarrow 1Ag^+ + 1Br^-$

x x

$K = x^2 = 5.4 \times 10^{-13}$

$x = 7.3 \times 10^{-7}$

72. The Ksp ~~solubility~~ of CaF_2 is 3.9×10^{-11} mol/L. What is the solubility ~~K_{sp}~~ of CaF_2?

a) 3.9×10^{-11}
b) 6.2×10^{-6}
c) 3.4×10^{-4}
[✓ d) 2.1×10^{-4}]
e) none of the above

$CaF_2(s) \rightarrow 1Ca^{2+} + 2F^-$

x $2x$

$K = (x)(2x)^2 = 3.9 \times 10^{-11}$

$4x^3 = 3.9 \times 10^{-11}$

$x^3 = 9.75 \times 10^{-12}$

$x = 2.1 \times 10^{-4}$

S
S
T
S

73. What is the solubility of barium sulfate in otherwise pure water? The K_{sp} value for barium sulfate is 1.1×10^{-10}.

a. 7.4×10^{-6} *M*
b. 5.5×10^{-11} *M*
[✓ c.] 1.0×10^{-5} *M*
d. 2.2×10^{-9} *M*
e. 1.1×10^{-10} *M*

$BaSO_4 \rightarrow Ba^{2+} + SO_4^{2-}$

$K = x^2$

74. Stalactites—the long, icicle-like formations that hang from the ceilings of caves—are formed from calcium carbonate. The K_{sp} of calcium carbonate is 4.5×10^{-9}. What is the **concentration of calcium ions** in a saturated calcium carbonate solution?

a. 0.00104 *M*
b. 4.5×10^{-9} *M*
[✓ c.] 6.7×10^{-5} *M*
d. 2.25×10^{-9} *M*
e. 4.5×10^{-5} *M*

$CaCO_3 \rightarrow \underset{x}{1Ca^{2+}} + \underset{x}{1CO_3^{2-}}$

$K = 4.5 \times 10^{-9} = x^2$

S
S
T
S

75. Lead pipes were used at one time for delivering drinking water. What is the maximum possible concentration of lead in this water if it comes from lead(II) hydroxide ($K_{sp} = 2.8 \times 10^{-16}$) dissolving from the surface of the pipes? Note the EPA limit on lead in drinking water is 7.2×10^{-8} *M*.

[✓ a.] 4.1×10^{-6} *M*
b. 1.6×10^{-8} *M*
c. 6.5×10^{-6} *M*
d. 5.1×10^{-6} *M*
e. 8.3×10^{-9} *M*

$Pb(OH)_2 \rightarrow \underset{x}{1Pb^{2+}} + \underset{2x}{2OH^-}$

$K = 2.8 \times 10^{-16} = 4x^3$

76. The solubility of $PbBr_2$ is 0.427 g per 100 mL of solution at 25°C. Determine the value of the solubility product constant for this strong electrolyte.

1. Find Ksp from solubility
2. Solubility given in g/mL rather than molarity
3. Here, needed to calculated mw in order to convert g => mol

a. 5.4×10^{-4}
b. 2.7×10^{-4}
c. 3.1×10^{-6}
d. 1.6×10^{-6}
[✓ e.] 6.3×10^{-6}

$x\,\frac{mol}{L} = \left(\frac{.427g}{.1L}\right)\left(\frac{1\,mol}{367g}\right)$

$= .0116\,\frac{mol}{L} = x$

$PbBr_2(s) \rightarrow \underset{x}{Pb^{2+}} + \underset{2x}{2Br^-}$

$K = (.0116)(.0233)^2$

$K = [Pb^{2+}][Br^-]^2$

77. Purveyors of salts from the Dead Sea advertise that it is healthy to bathe in a saturated solution of magnesium chloride ($MgCl_2$, 95.21 g/mol, $K_{sp} = 740$). How much magnesium chloride would you have to purchase to make up 10.0 L of bath water saturated with magnesium chloride?

1. Ksp => molar solubility
2. Actual volume => actual moles
3. Actual moles => actual grams

I
S

a. 9.0 kg
b. 12 kg
c. 57 kg
[✓ d.] 5.4 kg
e. 7.2 kg

$x\,g = \frac{10L}{} \left|\frac{5.698\,mol}{1L}\right|\frac{95.21g}{1\,mol}\left|\frac{1\,kg}{1000g}\right.$

$L \rightarrow mol \rightarrow g$

$MgCl_2 \rightarrow \underset{x}{Mg^{2+}} + \underset{2x}{2Cl^-}$

$4x^3 = 740$

$x = 5.698\,mol/L$

78. Stalactites—the long, icicle-like formations that hang from the ceilings of caves—are formed from calcium carbonate. Droplets saturated with calcium carbonate hang and evaporates, leaving solid calcium carbonate behind. The K_{sp} of calcium carbonate is 4.5×10^{-9}. What is the volume of water droplets saturated with calcium carbonate that would be required to form a small stalactite that had a mass of 1.0 kg?

1. Ksp => molar solubility
2. Actual grams => actual moles
3. Actual moles => volume

$mass \xrightarrow{} volume$

$\boxed{molarity}$

a. 1.4×10^{6} L
[✓ b.] 1.5×10^{5} L
c. 4.5×10^{9} L
d. 4.5×10^{4} L
e. 1.5×10^{4} L

$CaCO_3 \rightarrow Ca^{2+} + CO_3^{2-}$

$K = (4.5 \times 10^{-9}) = x^2$

$\boxed{x = 6.71 \times 10^{-5}\,\frac{mol}{L}}$

$CaCO_3 = 100.\frac{g}{mol}$

$x\,L = 1.0\,kg \left|\frac{1000g}{1\,kg}\right|\frac{1\,mol}{100g}\left|\frac{1\,L}{6.71\times10^{-5}\,mol}\right. = \boxed{1.5\times10^{5}\,L}$

$kg \rightarrow g \rightarrow mol \rightarrow L$

Common ions:
1. Inhibit solubility
2. Make concentrations of cation and anion unequal
3. Simplify solubility calculations 4. Qualitative LeChatelier push to the left

Impact on Solubility When Common Ions are Present

79. What is the solubility of barium sulfate in a solution also containing 0.050 *M* sodium sulfate (which is fully soluble)? The K_{sp} value for barium sulfate is 1.1×10^{-10}.

a. 7.4×10^{-6} *M*
b. 5.5×10^{-11} *M*
c. 1.0×10^{-5} *M*
[✓ d.] 2.2×10^{-9} *M*
e. 1.1×10^{-10} *M*

$BaSO_4 \rightleftharpoons Ba^{2+} + SO_4^{2-}$
$x \quad x \quad \cancel{x} + .050$
x
$.050$

$K = 1.1 \times 10^{-10} = (x)(.050)$

Finding the solubility for the "other" ion when one is known.
The "other" ion defines the overall solid's solubility

80. What is the solubility (in moles/L) of CuCl in a solution that also contains 0.020 M $CuNO_3$ (the latter is fully soluble). (K_{sp} CuCl = 1.0×10^{-6})

a) 1.0×10^{-6}
b) 1.0×10^{-3}
[✓ c)] 5.0×10^{-5}
d) 2.7×10^{-7}
e) none of the above

$CuCl(s) \rightleftharpoons Cu^{+} + Cl^{-}$
$.020 \quad x$

$K = 1 \times 10^{-6} = (x)(.020)$

81. When sodium chloride is added to a saturated solution of lead(II) chloride, some of the lead(II) chloride precipitates. This results from what is called __________

[✓ a.] the common ion effect.
b. selective precipitation.
c. supersaturation.
d. a solubility anomaly.
e. deionization.

$PbCl_2(s) \rightleftharpoons Pb^{2+} + 2Cl^{-}$
$Cl^{\ominus}$ extra

82. Consider a saturated solution of AgCl in water. Compared to the original concentrations of the $[Ag^+]$ and $[Cl^-]$, how would the concentrations be different after some NaCl was added to the solution?

a. The resulting $[Ag^+]$ and $[Cl^-]$ would both end up higher than they were originally.
b. The resulting $[Ag^+]$ and $[Cl^-]$ would both end up lower than they were originally.
c. The resulting $[Ag^+]$ would be larger, but the resulting $[Cl^-]$ would be smaller.
[✓ d.] The resulting $[Ag^+]$ would be smaller, but the $[Cl^-]$ would be larger.
e. $[Ag^+]$ and $[Cl^-]$ would remain the same because the solution is saturated.

$AgCl \rightleftharpoons Ag^{+} + Cl^{-}$
$Cl^{\ominus}$ extra

83. Which of the following, when added to a saturated solution of LaF_3, will increase the solubility LaF_3? (K_{sp}=4×10^{-17})

a) Addition of soluble KNO_3
b) Addition of soluble ~~$La(NO_3)_3$~~
c) Addition of soluble ~~KF~~
[✓ d)] Addition of HNO_3

$LaF_3 \rightleftharpoons La^{3+} + F^{-}$
$\downarrow H^{\oplus}$
HF

84. When a solid partially dissolves to produce two ions, the solubility can be reduced if either of the two ions is supplied by a different source. This reduced concentration of one of the ions, and reduced solubilty of the parent ionic solid, is the result of what's called the __________

a. ionic suppression effect.
b. ~~counter ion effect.~~
[✓ c.] common ion effect.
d. excession effect.
e. supersaturation effect.

1. Any solid that produces a BASIC anion will be more soluble at acidic pH
2. Any solid that produces a NONBASIC anion won't care about pH.
3. Acid reacts with and removes basic anions, so LeChatelier "pulls" to the right
4. For a metal hydroxide, the hydroxide concentration can be quantified from pH, so solving for the "other" ion become easy

Impact of pH on Solubility

S S T S

85. Which of the following statements is **FALSE**?

- a. Any solid that produces a basic anion will become more soluble at low pH T
- b. Low pH increases solubility for solids that produce basic anions, because the acid reduces the concentration of the basic anion. LeChatelier then forces more solid to dissolve, resulting in elevated concentration of the cation. T
- c. For solids that produce basic anions, the concentration of anion is lower than the concentration of the cation at low pH
- d. For metal hydroxides, solubility decreases at high pH where hydroxide concentration is high
- [✓ e.] For solids with nonbasic anions like chlorides or bromides, solubility is still higher at low pH.

86. Manganese carbonate ($MnCO_3$) has low solubility in neutral water (8 mg/L), but dissolves completely in water whose pH is low. Which of the following explanations are true?

1. At low pH, H^+ continuously and irreversibly removes carbonate ion T
2. At low pH, OH^- continuously removes Mn^{2+} ion F
3. All solids always become completely soluble in acidic solution F

$CO_3^{2-} + 2H^+ \rightarrow H_2CO_3$

$MnCO_3 \overset{\longrightarrow}{\longrightarrow} Mn^{2+} + CO_3^{-2}$

- [✓ a) 1 only]
- b) 2 only
- c) 3 only
- d) 1 and 3 only
- e) 2 and 3 only

S S T S

87. As the pH decreases, the solubility of __________ would increase.

- a. lead(II) chloride
- b. silver(I) iodide
- [✓ c.] calcium carbonate
- d. mercury(I) bromide
- e. silver(I) chloride

basic anion

88. Which of the following compounds would *not* have a pH dependent solubility?

- a. $Mg(OH)_2$
- b. PbS
- [✓ c.] AgI ← nonbasic anion
- d. Na_2O
- e. PbS

Nonbasic
Cl^-
N_3^-
I^-
NO_3^-
ClO_4^-
HSO_4^-
} strong acid conj

S S T S

89. Which of the following compounds would *not* have a pH dependent solubility?

- a. $Fe(OH)_2$
- b. $Mn(N_3)_2$
- c. AgF
- [✓ d.] $PbBr_2$ ← nonbasic anion
- e. AgCN

90. Which of the following compounds would *not* have a pH dependent solubility?

- a. $FeCO_3$
- b. $Zn(OH)_2$
- c. $Ca(HCO_3)_2$
- d. $FePO_4$
- [✓ e.] Hg_2Cl_2

General Chemistry II Jasperse ANSWERS
Buffer/Titration/Solubility. Extra Practice Problems

1. C
2. **Answer:** A buffer consists of a weak acid and its conjugate base in roughly equal amounts. If acid is added to the solution, it is consumed by the conjugate base. If base is added to the solution, it is consumed by the weak acid. If the amounts are such that the ratio of conjugate base/weak acid concentrations doesn't change much, then the pH doesn't change much. Dilution does not affect the pH because this concentration ratio doesn't change upon dilution.
3. B
4. B
5. C
6. A
7. A
8. D
9. A
10. D
11. B
12. C
13. C
14. B
15. D
16. E
17. D
18. C
19. C
20. 9.43
21. D
22. 3.08
23. 3.61
24. B
25. C
26. A
27. A
28. C
29. A
30. D
31. D
32. D
33. A
34. C
35. A
36. 2.74
37. 3.16
38. 9.03
39. 9.55
40. A
41. 9.07g
42. D
43. B
44. D
45. B
46. D
47. B
48. D
49. C
50. B
51. B
52. B
53. D
54. A
55. A
56. B
57. B
58. D
59. B
60. A
61. B
62. D
63. E
64. A
65. B
66. C
67. $PbCl_2(s) \rightarrow Pb^{2+}(aq) + 2Cl^{-}(aq)$

    $K_{sp} = [Pb^{2+}]\,[Cl^{-}]^2$
68. B
69. D
70. D
71. B
72. D
73. C
74. C
75. A
76. E
77. D
78. B
79. D
80. C
81. A
82. D
83. D
84. C
85. E
86. A
87. C
88. C
89. D
90. E