

Recitation Worksheet Twelve

Name:

Key

MyID:

Textbook:

Chemistry & Chemical Reactivity

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Instructions:

- This recitation worksheet covers Ch. 16.5, 17.1-17.2
- Please enter your first and last name as it appears on the eLC roster (do not use a nickname that is not reflected in eLC).
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 - You must make sure the pages are in the correct order and have the same layout as the original worksheet when submitting to Gradescope regardless of your submission type.
 - Answers must be written in the corresponding answer boxes.
 - You must show your work when appropriate.
- This worksheet is due no later than **12:00 PM (noon) on the Saturday, November 16th.**
- A periodic table and formula sheet are attached to the end of this worksheet. Please keep these attached to your worksheet in the correct order when submitting to Gradescope.

A buffer solution is a solution of a weak acid & salt of its conjugate base or a weak base & the salt of its conjugate acid.

1. Which of the aqueous solutions below are buffer solutions? Select all that apply. Insert letters without spaces in the answer box, example ABCD.

CEF

- A. 0.100 M KBr
 B. 0.200 M NaCl and 0.200 M NH_4Cl
 C. 0.100 M CH_3NH_2 and 0.150 M $\text{CH}_3\text{NH}_3^+\text{Cl}^-$
 D. 0.100 M HCl and 0.050 M NaNO_2
 E. 0.100 M HCl and 0.200 M NaCH_3COO
 F. 0.100 M Na_2HPO_4 and 0.100 M NaH_2PO_4
 G. 0.100 M CH_3COOH and 0.100 M $\text{NaCH}_2\text{CH}_2\text{COO}$

A. KBr is a pH-neutral salt.

B. NaCl is a pH-neutral salt & NH_4^+Cl^- is an acidic salt.

C. CH_3NH_2 is a weak base & $\text{CH}_3\text{NH}_3^+\text{Cl}^-$ the salt of its conjugate acid.

Buffer mixture

D. $\text{HCl} + \text{NaNO}_2 \rightarrow \text{HNO}_2 + \text{NaCl}$

I	0.100 mol	0.050 mol	0	0
C	-0.050 mol	0.050 mol	+0.050	+0.050
F	0.050 mol	0 mol	0.050 mol	0.050 mol

From the ICF table HCl (strong acid), HNO_2 (weak acid), & NaCl (pH-neutral salt) which none of these components form a buffer.

E. $\text{HCl} + \text{CH}_3\text{COONa} \rightarrow \text{CH}_3\text{COOH} + \text{NaCl}$

I	0.100 mol	0.200 mol	0	0
C	-0.100	-0.100 mol	+0.100 mol	+0.100 mol
F	0	0.100 mol	0.100 mol	0.100 mol

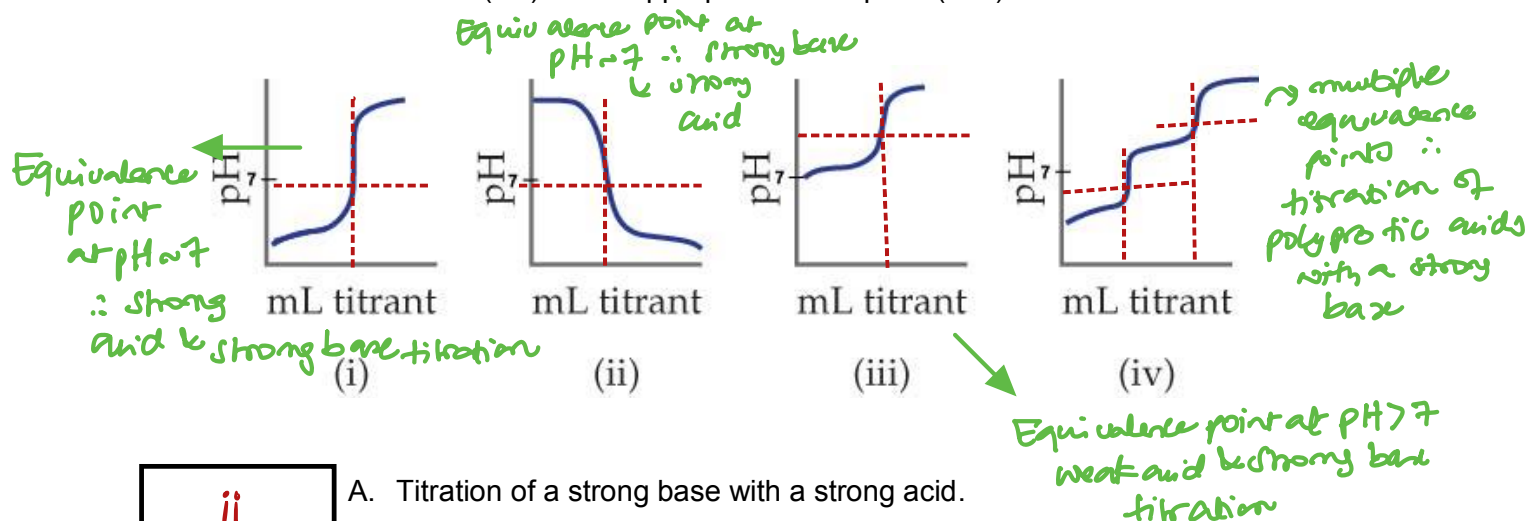
Buffer mixture

F. NaH_2PO_4 & Na_2HPO_4 are both basic salts but $\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$ are conjugates.

acid conjugate base
 Buffer mixture

G. Although CH_3COOH is a weak acid & $\text{NaCH}_2\text{CH}_2\text{COO}$ is a weak base they are not conjugates.

2. Match the titrations curves (i-iv) to the appropriate description (A-E).



ii

A. Titration of a strong base with a strong acid.

iii

B. Titration of a weak acid with a strong base.

i

C. Titration of a strong acid with a strong base.

iv

D. Titration of a polyprotic acid with a strong base.

$$K_a \text{ of } \text{NH}_4^+ = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

$$\therefore \text{p}K_a = -\log(5.6 \times 10^{-10}) = 9.26$$

3. In the laboratory, you were asked to prepare a buffer solution with $\text{pH} = 10.50$. How many **grams** of NH_4Cl (molar mass = 53.5 g/mol) would you add to 625 mL of 0.258 M NH_3 to prepare a buffer with $\text{pH} = 10.50$? Assume that the solution's volume remains constant. K_b of $\text{NH}_3 = 1.8 \times 10^{-5}$.

0.49

① Using Henderson-Hasselbalch equation to determine the ratio of base to acid

$$\text{pH} = \text{p}K_a + \log \frac{[\text{NH}_3]}{[\text{NH}_4^+]} \Rightarrow 10.50 = 9.26 + \log \frac{0.258}{[\text{NH}_4^+]}$$

(2 decimal places)

$$\therefore 1.25 = \log \frac{0.258}{[\text{NH}_4^+]} \rightarrow \text{taking the inverse of log on both sides}$$

$$17.56820923 = \frac{0.258}{[\text{NH}_4^+]} \Rightarrow \therefore [\text{NH}_4^+] = 0.014685617 \text{ M}$$

② Use the molarity of NH_4Cl to determine the grams of NH_4Cl added to 625 mL of 0.258 M NH_3

$$625 \text{ mL} \times \frac{0.014685617 \text{ mol}}{1000 \text{ mL}} \times \frac{1 \text{ mol } \text{NH}_4\text{Cl}}{1 \text{ mol } \text{NH}_4^+} \times \frac{53.5 \text{ g } \text{NH}_4\text{Cl}}{1 \text{ mol } \text{NH}_4\text{Cl}}$$

$$= 0.49 \text{ g}$$

4. If 0.100 M of the solutions below were provided to you, which **two solutions** would you use to prepare a buffer with pH = 3.50? **Select all that apply.** Insert letters without spaces in the answer box, example **AB**.

AE

- A. Formic acid (HCOOH , $\text{pK}_a = 3.74$)
 B. Acetic acid, (CH_3COOH , $\text{pK}_a = 4.74$)
 C. Phosphoric acid (H_3PO_4 , $\text{pK}_{a1} = 2.15$)
 D. Sodium acetate (NaCH_3COOH)
 E. Sodium formate (NaHCOO)
 F. Sodium dihydrogen phosphate (NaH_2PO_4)

- to prepare a buffer with $\text{pH} = 3.50$ you need to choose an acid with a pK_a close to the pH of the buffer & in this case this is formic acid
 - the conjugate base salt of formic acid is sodium formate ($\text{HCOOH}/\text{HCOONa}$)

5. Calculate the **final pH** in each of the **titration scenarios** below:

- A. The titration of 25.00 mL of 0.160 M HCl with 15.00 mL of 0.242 M NaOH .

strong acid

strong base

2.034

①



I	0.00400	0.00363	0	-
C	-0.00363	-0.00363	+0.00363	+0.00363
F	0.00037	0	0.00363	0.00363

left over HCl

→ pH-neutral salt $\therefore \text{Na}^+$ & Cl^- ions are spectator ions

- ② Calculate the new molarity of HCl .

$$\frac{0.00037 \text{ mol}}{25.00 + 15.00 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.00925 \text{ M}$$

- ③ pH of solution will be the pH of the remaining HCl
 $\therefore \text{pH} = -\log [0.00925]$
 $= 2.034$

8 decimal places

strong base

- B. The titration of 25.00 mL of 0.100 M CH_3COOH (K_a of $\text{CH}_3\text{COOH} = 1.7 \times 10^{-5}$) with 12.5 mL of 0.200 M NaOH .

weak acid

8.797

①



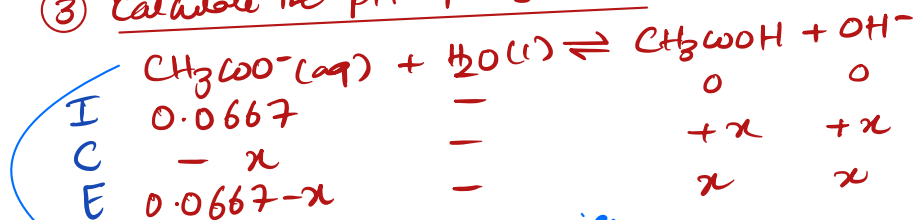
I	0.00250 mol	0.00250	0	-
C	-0.00250 mol	-0.00250	+0.00250 mol	-
F	0	0	0.00250 mol	-

↓
 salt with basic pH

② Molarity of CH_3COONa

$$\frac{0.00280 \text{ mol}}{25.00 + 12.5 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.0667 \text{ M}$$

③ Calculate the pH of $\text{CH}_3\text{COO}^-\text{Na}^+$ (Na^+ is pH-neutral)



solving for a base \therefore determine

$$K_b \therefore K_b = \frac{1.0 \times 10^{-14}}{1.7 \times 10^{-5}} = 5.9 \times 10^{-10}$$

$$\therefore 5.9 \times 10^{-10} = \frac{[x][x]}{[0.0667 - x]}$$

$$x^2 = 5.9 \times 10^{-10} \times 0.0667$$

↓
take square root
of both sides

$$x = \pm 6.2638 \times 10^{-6}$$

$$\textcircled{4} \text{ pOH} = -\log [6.2638 \times 10^{-6}]$$

$$= 5.203$$

$$\therefore \text{pH} + \text{pOH} = 14.00$$

$$\therefore \text{pH} = 14.00 - 5.20$$

$$= 8.797$$

6. What is the pH of a mixture of 0.012 M of $\text{C}_6\text{H}_5\text{COOH}$ ($K_a = 6.3 \times 10^{-5}$) and 0.033 M $\text{NaC}_6\text{H}_5\text{COO}$?

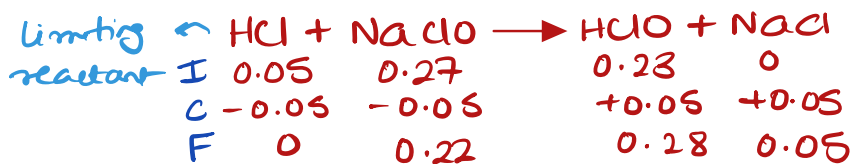
4.64

$$\begin{aligned} \text{pH} &= \text{p}K_a + \log \frac{[\text{NaC}_6\text{H}_5\text{COO}]}{[\text{C}_6\text{H}_5\text{COOH}]} \quad \text{Henderson-Hasselbalch equation} \\ &= -\log(6.3 \times 10^{-5}) + \log \frac{[0.033]}{[0.012]} \\ &= 4.63999 \\ &\sim 4.64 \end{aligned}$$

7. A solution is prepared by dissolving 0.23 mol of hypochlorous acid and 0.27 mol of sodium hypochlorite in water sufficient to yield 1.00 L of solution. The addition of 0.05 mol of HCl to this buffer solution causes the pH to drop slightly. The pH does not decrease drastically because the HCl reacts with the _____ present in the buffer solution. The K_a of hypochlorous acid is 1.36×10^{-3} .

C

- A. H_2O
B. H_3O^+
C. Hypochlorite ion
D. Hypochlorous acid
E. This is a buffer solution. The pH does not change upon addition of acid or base.

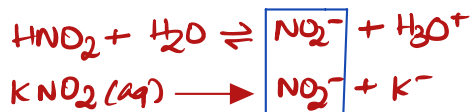


8. In which of the solutions given below would nitrous acid (HNO_2) ionize less than it does in pure water?

C

- A. NaCl
B. KNO_3
C. KNO_2
D. NaClO_4

Common ion effect



Common ion \therefore shifts the equilibrium position of HNO_2 to the left \therefore HNO_2 ionizes less

$$\text{Moles of CH}_3\text{COOH} = 0.300 \text{ L} \times \frac{0.250 \text{ mol}}{1 \text{ L}} = 0.075 \text{ mol}$$

Buffer

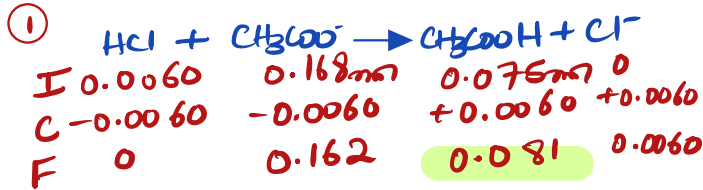
$$\text{Moles of NaCH}_3\text{COO} = 0.300 \text{ L} \times \frac{0.560 \text{ mol}}{1 \text{ L}} = 0.168 \text{ mol}$$

9. To a 0.300 L buffer solution consisting of 0.250 M CH_3COOH and 0.560 M NaCH_3COO , 0.0060 mol HCl is added. What are the moles and concentration of CH_3COOH after the addition of HCl? Assume that the volume of the buffer does not change upon the addition of HCl.

Strong acid

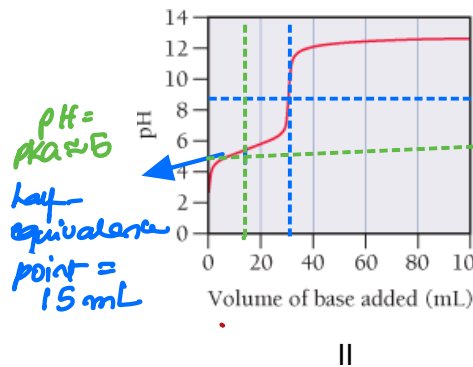
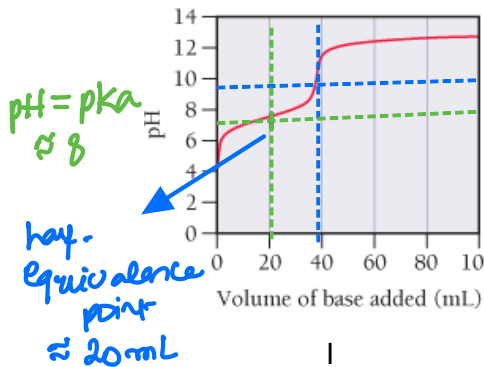
should react with base component in the buffer (CH_3COO^-)

- C**
- A. 0.0060 mol, 0.020 M
 - B. 0.162 mol, 0.54 M
 - C. 0.081 mol, 0.27 M
 - D. 0.075 mol, 0.250 M
 - E. 0.168 mol, 0.560 M



② Moles of $\text{CH}_3\text{COOH} = \frac{0.081 \text{ mol}}{0.300 \text{ L}} = 0.27 \text{ M}$

10. You are provided with the titration curves I and II for two weak acids titrated with 0.100 M NaOH. Refer to the titration curves to answer the following questions:



- A. Which acid is more concentrated? **I**
- more concentrated acid will produce a greater $[\text{H}_3\text{O}^+]$ & will consume greater moles of NaOH
 \therefore greater volume of NaOH

- B. Which acid has the larger K_a ? **II**
- $\text{pH} = \text{pKa}$ at half equivalence point
 the \downarrow the pKa the larger the K_a & the stronger the acid

At half-equivalence point half the moles of the acid are converted to moles of conjugate salt
 $\therefore \text{pH} = \text{pKa} + \log \frac{[\text{A}^-]}{[\text{HA}]}$ if moles are the same $\therefore \text{pH} = \text{pKa}$

11. Which solution has the **greatest** buffering capacity?

C

- A. 0.335M $\text{HC}_2\text{H}_3\text{O}_2$ and 0.497 M $\text{NaC}_2\text{H}_3\text{O}_2$
- B. 0.520 M $\text{HC}_2\text{H}_3\text{O}_2$ and 0.116 M $\text{NaC}_2\text{H}_3\text{O}_2$
- C. 0.820 M $\text{HC}_2\text{H}_3\text{O}_2$ and 0.715 M $\text{NaC}_2\text{H}_3\text{O}_2$
- D. 0.120 M $\text{HC}_2\text{H}_3\text{O}_2$ and 0.115 M $\text{NaC}_2\text{H}_3\text{O}_2$

the greatest buffer capacity is when the buffer components are large & equal

12. A strong base such as KOH is mixed in a specific proportion with the weak acid HA to make a buffer. Which of the diagrams below is a correct representation of the buffer solution?

A



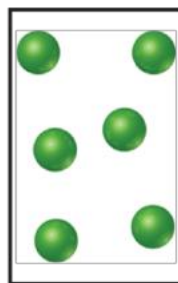
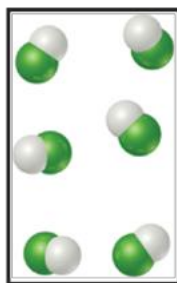
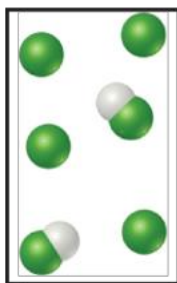
= HA



= A^-



= OH^-



A

B

C

D



the weak acid & its conjugate base salt should be present in order for the solution to be a buffer solution

13. What change will be caused by addition of a small amount of HCl to a solution containing fluoride ions and hydrogen fluoride?

Strong acid

HF/F⁻

D

- A. The concentration of hydronium ions will increase significantly.
- B. The concentration of fluoride ions will increase as will the concentration of hydronium ions. *conc. of F⁻ ↓ & H₃O⁺ ↑*
- C. The concentration of hydrogen fluoride will decrease, and the concentration of fluoride ions will increase. *HF ↑ & F⁻ ↓*
- D. The concentration of fluoride ion will decrease, and the concentration of hydrogen fluoride will increase.
- E. The fluoride ions will precipitate out of solution as its acid salt.



more HF is produced ∴

the pH of the buffer will increase

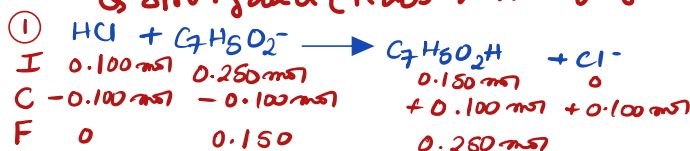
Buffer

14. A 1.00 L buffer solution is 0.150 M in HC₇H₅O₂ and 0.250 M in LiC₇H₅O₂. Calculate the pH of the solution after the addition of 100.0 mL of 1.00 M HCl. The K_a for HC₇H₅O₂ is 6.5 × 10⁻⁵.

E

- A. 4.19
- B. 5.03
- C. 4.41
- D. 3.34
- E. 3.97

↳ Strong acid (reacts with any weak base)



② Using Henderson-Hasselbalch equation to calculate the pH of the buffer

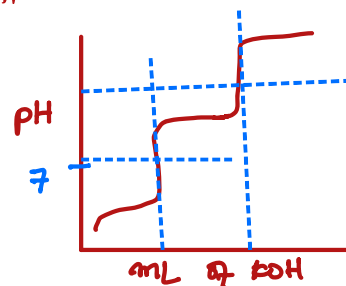
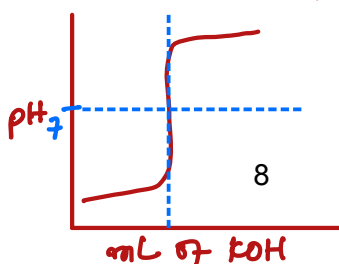
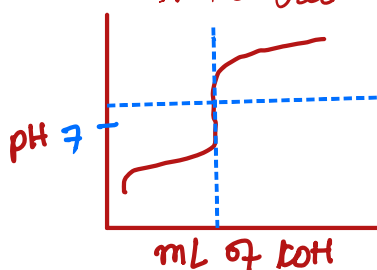
$$\begin{aligned} \text{pH} &= \text{pK}_a + \log \frac{[\text{C}_7\text{H}_5\text{O}_2^-]}{[\text{HC}_7\text{H}_5\text{O}_2]} \\ &= -\log(6.5 \times 10^{-5}) + \log \frac{[0.150]}{[0.250]} = 3.96523 \sim 3.97 \end{aligned}$$

15. You are working in the lab with three acidic solutions. **Solution 1** is 0.1 M of a weak monoprotic acid, **solution 2** is 0.1 M of a strong monoprotic acid and **solution 3** is a 0.1 M weak diprotic acid. Each of the former solutions has been titrated with a 0.2 M KOH solution. Which quantity is the same for all the three solutions?

B

- A. The volume required to reach the final equivalence point. *diprotic acid has two equivalence points*
- B. The volume required to reach the first equivalence point.
- C. The pH at the first equivalence point. *pH at first equivalence point for strong acid is ≈ 7 for weak acids is > 7*
- D. The pH at one-half the first equivalence point. *not the same*
- E. None of the quantities is the same for the three solutions.

At equivalence point, the mol of the acid = mol of the base. Since all the acids have the same conc. (0.1 mol/L) & they are titrated with 0.2 M KOH ∴ the volume to reach the first equivalence point



16. Which of the following statements is accurately describes the common-ion effect? *decreased*

C

- A. The extent of ionization of a weak electrolyte is increased by adding to the solution a strong electrolyte that has an ion in common with the weak electrolyte. *decreased*
- B. The solubility of a slightly soluble salt is increased by the presence of a second solute that provides a common ion to the system. *decreased*
- C. The common ion effect occurs when a solubility equilibrium is shifted due to a second compound that contains an ion common with the first.
- D. The common ion effect is that common ions precipitate all counter-ions.
- E. None of the above statements accurately describe the common-ion effect.

choice C is the definition of the common-ion effect

17. 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 $\text{Na}_2\text{HPO}_4/\text{NaH}_2\text{PO}_4$ was needed?

C

- A. 1.2/1.0
B. 1.0/1.0
C. 1.6/1.0
D. 0.96/1.0
E. 0.90/1.0

$$\text{pH} = \text{pK}_a + \log \frac{[\text{Base}]}{[\text{Acid}]}$$

Buffer
conjugate base

Acid

$$7.4 = -\log(6.2 \times 10^{-8}) + \log \frac{[\text{Na}_2\text{HPO}_4]}{[\text{NaH}_2\text{PO}_4]}$$

$$7.4 = 7.207608811 + \log \frac{[\text{Na}_2\text{HPO}_4]}{[\text{NaH}_2\text{PO}_4]}$$

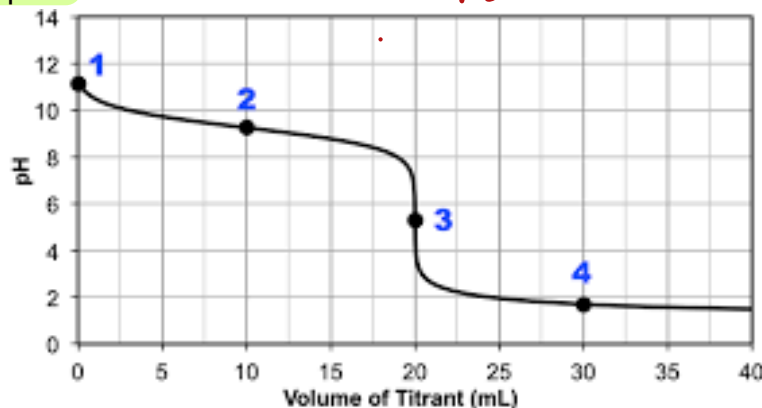
$$0.192391689 = \log \frac{[\text{Na}_2\text{HPO}_4]}{[\text{NaH}_2\text{PO}_4]}$$

$$\therefore \frac{[\text{Na}_2\text{HPO}_4]}{[\text{NaH}_2\text{PO}_4]} = 10^{0.19239168} = 1.557369586 \sim \frac{1.6}{1.0}$$

18. At what point in this titration curve is the pH = pK_a?

B

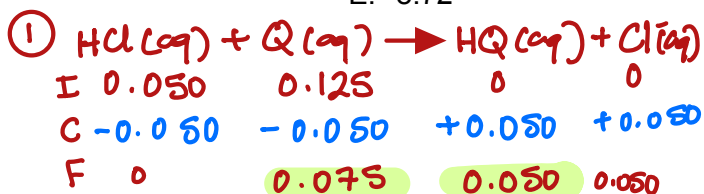
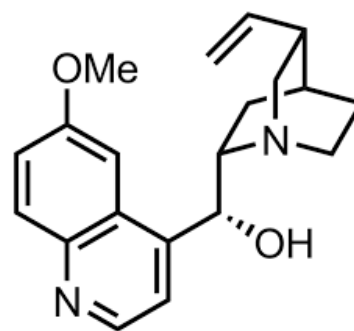
- A. 1
B. 2
C. 3
D. 4
E. Not enough information to determine



- the titration curve above represents the titration of a weak base (starting pH ~ 11) with a strong acid (pH ~ 1.8)
- pH = pK_a at half equivalence point (2)

19. Quinine is a weak base that is used to treat malaria which is caused by the parasite Plasmodium falciparum. A 25.0 mL solution of quinine was titrated with 1.00 M hydrochloric acid. 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 $pK_b = 5.10$.

- D**
- A. 5.10
B. 4.92
C. 8.90
D. 9.08
E. 8.72



Buffers

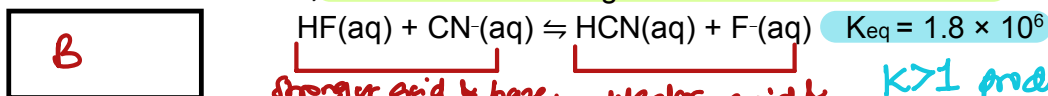
② $pH = pK_a + \log \frac{[base]}{[acid]}$

\downarrow

$pK_a = 14.00 - pK_b = 14.00 - 5.10 = 8.90$

$\therefore pH = 8.90 + \log \frac{[0.075]}{[0.050]} = 9.076091259 \sim 9.08$

20. For the reaction shown, which of the following statements would be false?



- A. HF is the strongest acid **True**
- B. Fluoride anion is the strongest base **False**
- C. Cyanide anion is the strongest base **True**
- D. The solution will contain more HCN than HF at equilibrium

\downarrow

True since $K > 1$ (product favored)

$K > 1$ product favored

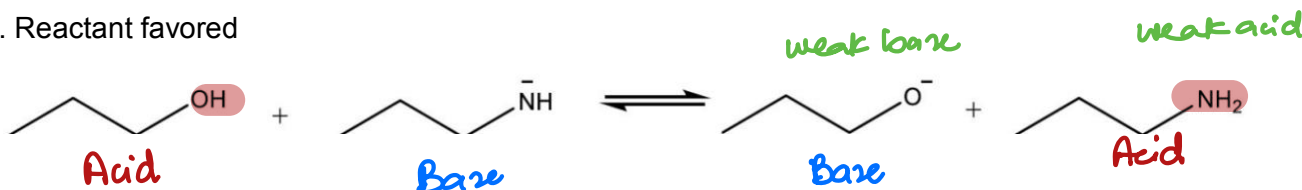
(An acid-base reaction equilibrium always favors the side of the weaker acid/base)

21. On which side of the equation does the equilibrium lie for this reaction?

A

- A. Product favored
B. Reactant favored

↓
Equilibrium will favor the weaker acid / weaker base



Let's compare the acids:

on the left side the acidic hydrogen is attached to oxygen while on the right side the acid hydrogen is attached to nitrogen. The stronger acid will have the acidic hydrogen attached to the more electronegative atom ($O > N$ in EN) the acidic hydrogen attached to the more electronegative atom.

You can also compare bases: since $N < O$ in EN it cannot handle the negative charge as well as O can. \therefore CCCC[NH-] is a stronger base.

22. What is the pH of an aqueous solution that contains 0.226 M potassium hydrogen phosphate, K_2HPO_4 , and 0.451 M potassium dihydrogen phosphate, KH_2PO_4 ?

(K_a for $H_3PO_4 = 7.52 \times 10^{-3}$, K_a for $H_2PO_4^- = 6.23 \times 10^{-8}$, K_a for $HPO_4^{2-} = 4.8 \times 10^{-13}$)

Buffer

C

- A. 1.824
B. 6.562
C. 6.905
D. 7.208
E. 12.019

$$pH = pK_a + \log \frac{[HPO_4^{2-}]}{[H_2PO_4^-]}$$
 (Note: $[HPO_4^{2-}]$ is labeled "From K_2HPO_4 " in blue)

$$pH = -\log(6.23 \times 10^{-8}) + \log \frac{[0.226]}{[0.451]}$$

$$= 6.905443851$$

$$\approx 6.905$$

Extra Practice Questions: these questions will not be graded

1. How many grams of sodium benzoate ($\text{NaC}_6\text{H}_5\text{COO}$, MW: 144.11) would need to be added to 250. mL of a 0.200 M benzoic acid ($\text{C}_6\text{H}_5\text{COOH}$) solution to achieve a pH of 3.98? The pK_a of benzoic acid is 4.20.

4.3 g

$$\textcircled{1} \text{ pH} = \text{pK}_a + \log \frac{[\text{C}_6\text{H}_5\text{COO}^-]}{[\text{C}_6\text{H}_5\text{COOH}]}$$

base *acid*

$$3.98 = 4.20 + \log \frac{[\text{C}_6\text{H}_5\text{COO}^-]}{[0.200]}$$

2 decimal places

$$-0.22 = \log \frac{[\text{C}_6\text{H}_5\text{COO}^-]}{[0.200]}$$

2 sig figs

$$10^{-0.22} = \frac{[\text{C}_6\text{H}_5\text{COO}^-]}{[0.200]}$$

$$0.6025595861 = \frac{[\text{C}_6\text{H}_5\text{COO}^-]}{[0.200]}$$

$$\therefore [\text{C}_6\text{H}_5\text{COO}^-] = 0.1205119172 \text{ M}$$

$$\textcircled{2} \frac{0.1205119172 \text{ mol } [\text{C}_6\text{H}_5\text{COO}^-]}{1 \text{ L}} \times \frac{1 \text{ mol } \text{NaC}_6\text{H}_5\text{COO}}{1 \text{ mol } \text{C}_6\text{H}_5\text{COO}^-} \times \frac{144.11 \text{ g } \text{NaC}_6\text{H}_5\text{COO}}{1 \text{ mol } \text{NaC}_6\text{H}_5\text{COO}} \times 0.250 \text{ L}$$

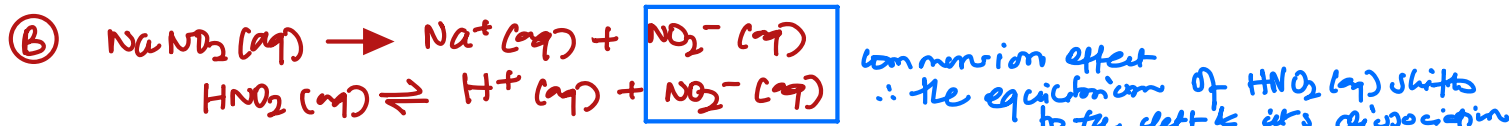
$$= 4.341743098 \text{ g} \approx 4.3 \text{ g}$$

2. Which of these statements regarding a solution of HNO_2 is false?

D

- A. Addition of NaOH will increase dissociation of HNO_2 , and the pH of the final solution will be higher $\text{NaOH(aq)} + \text{HNO}_2(\text{aq}) \rightarrow \text{NaNO}_2(\text{aq}) + \text{H}_2\text{O(l)}$
- B. Addition of NaNO_2 will decrease dissociation of HNO_2 , and the pH of the final solution will be higher
- C. Addition of HNO_3 will decrease dissociation of HNO_2 , and the pH of the final solution will be lower *A strong acid will decrease the dissociation of a weak acid*
- D. Addition of NaNO_2 will decrease dissociation of HNO_2 , and the pH of the final solution will be lower *higher*

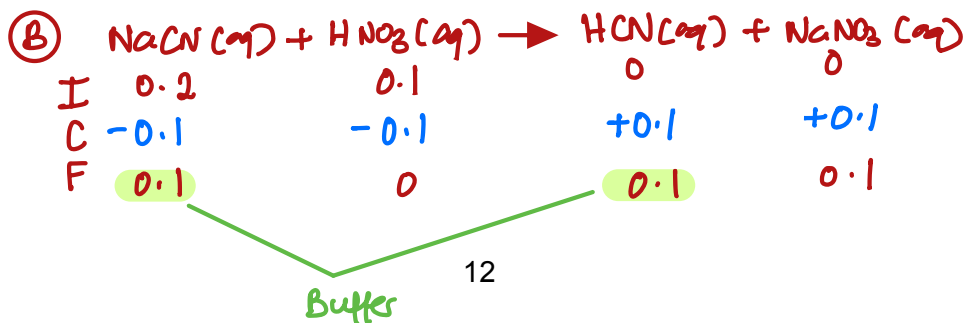
\textcircled{A} HNO_2 is neutralized \therefore by NaOH \therefore pH will increase (pH of final solution is higher)



3. Which of these would not result in a buffer solution if 100 mL of each solution is mixed?

C

- A. 0.10 M NaCH_3COO and 0.15 M CH_3COOH *weak base conjugate acid both are conjugate \therefore buffer \therefore pH will be high*
- B. 0.2 M NaCN and 0.1 M HNO_3
- C. 0.15 M HClO_4 and 0.12 M NaClO_4 *strong acid & its conjugate cannot be a buffer*
- D. 1.2 M $\text{C}_5\text{H}_5\text{N}$ and 0.9 M $\text{C}_5\text{H}_5\text{NHCl}$ *weak base conjugate acid \therefore buffer*
- E. 0.7 M NaOH and 1.2 M NaH_2PO_4





I 0.7

1.2

0

0

C - 0.7

-0.7

+0.7

-

F 0

0.6

0.7

Buffer

4. The following compounds are available as 0.10 M aqueous solutions: pyridine ($pK_b = 8.82$), triethylamine ($pK_b = 3.25$), HClO_4 , phenol ($pK_a = 9.96$), HClO ($pK_a = 7.54$), NH_3 ($pK_b = 4.74$) and NaOH . Identify two solutions that could be used to prepare a buffer with a pH of approximately 5.

A

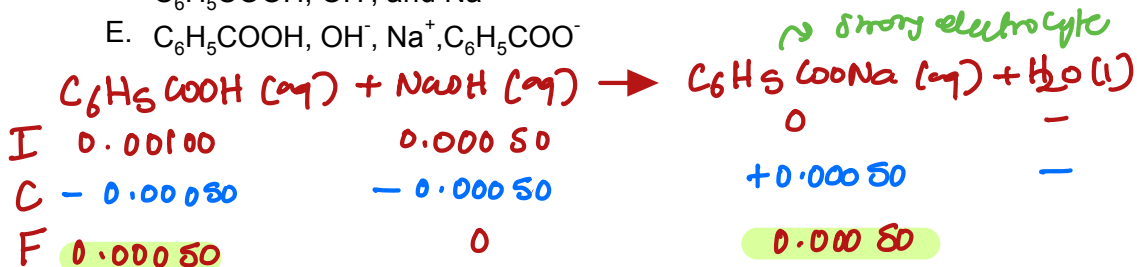
- A. pyridine and HClO_4
 B. triethylamine and HClO_4
 C. phenol and NaOH
 D. HClO and NaOH
 E. HClO and NH_3

$pK_a + pK_b = 14.00$
 $\therefore pK_a = 14.00 - pK_b$
 To prepare a buffer with the best buffer capacity, choose the compound with pK_a closest to $\text{pH} = 5$

5. In the titration of 50.0 mL of 0.0200 M $\text{C}_6\text{H}_5\text{COOH}(\text{aq})$ with 0.100 M $\text{NaOH}(\text{aq})$, what is/are the major species in the solution after the addition of 5.0 mL of $\text{NaOH}(\text{aq})$?

A

- A. $\text{C}_6\text{H}_5\text{COOH}$, $\text{C}_6\text{H}_5\text{COO}^-$, and Na^+
 B. $\text{C}_6\text{H}_5\text{COOH}$
 C. $\text{C}_6\text{H}_5\text{COO}^-$ and Na^+
 D. $\text{C}_6\text{H}_5\text{COOH}$, OH^- , and Na^+
 E. $\text{C}_6\text{H}_5\text{COOH}$, OH^- , Na^+ , $\text{C}_6\text{H}_5\text{COO}^-$



6. For the following titration, determine whether the solution at the equivalence point is acidic, basic or neutral and why:

A

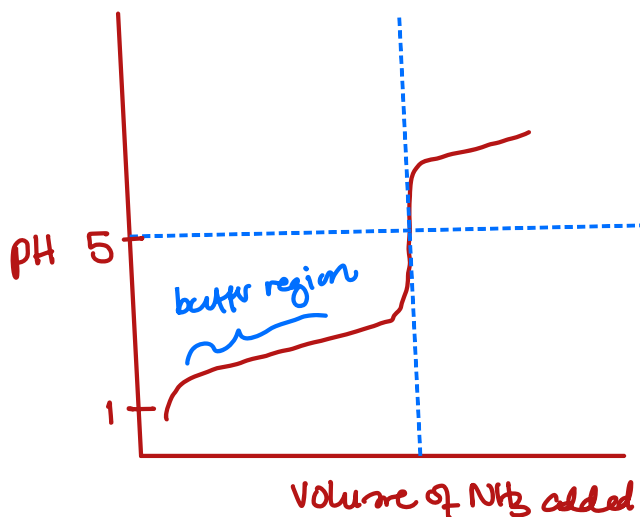
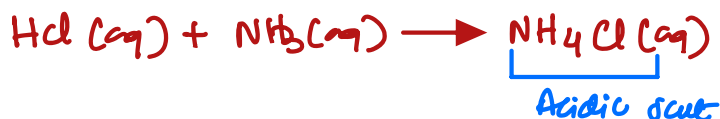
Strong acid

$\text{HCl}(\text{aq})$ is titrated with $\text{NH}_3(\text{aq})$

Weak base

\hookrightarrow moles of acid = moles of base

- A. acidic because of hydrolysis of NH_4^+
 B. basic because of hydrolysis of NH_3
 C. acidic because of hydrolysis of Cl^-
 D. acidic because of hydrolysis of HCl
 E. neutral salt of strong acid and strong base

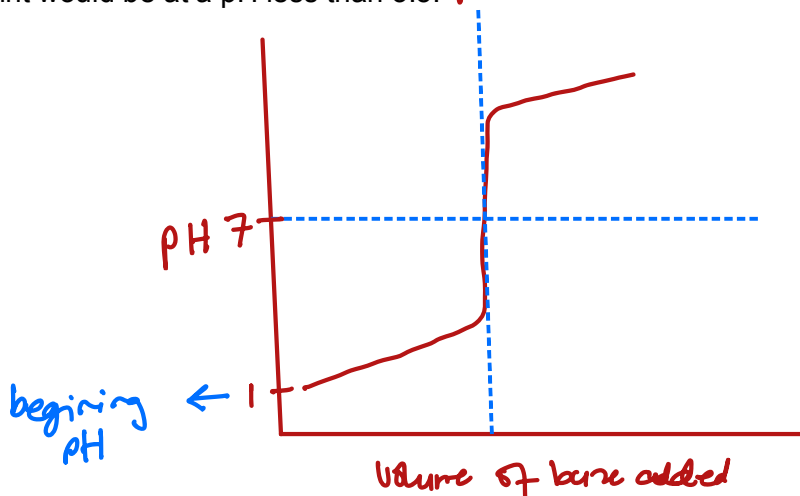


7. Which of the following statements correctly describe a typical titration curve for the titration of a strong acid by a strong base?

E

- I) The beginning pH is low. *True*
- II) The pH change is slow until near the equivalence point. *True*
- III) At the equivalence point, pH changes by a large value. *True*
- IV) Beyond the equivalence point, pH rises rapidly. *False*
- V) The equivalence point would be at a pH less than 3.5. *pH=7 False*

- A. I), III) and V)
- B. II), III) and IV)
- C. I), III) and IV)
- D. III), IV) and V)
- E. I), II) and III)



8. Choose the expression that gives the molar concentration of a $\text{H}_2\text{SO}_4(\text{aq})$ solution if 24.3 mL of a 0.105 M NaOH(aq) solution is required to titrate 60 mL of the acid.

D

- A. $(60 \times 24.3)/0.105 = 1.288571429 \times 10^4$
- B. $(60 \times 2)/(24.3 \times 0.105) = 47.03115814$
- C. $(24.3 \times 0.105)/(60) = 0.042526$
- D. $(24.3 \times 0.105)/(60 \times 2) = 0.0212625$
- E. $(60 \times 0.105)/(2 \times 24.3) = 0.129629296$



$$\frac{24.3 \text{ mL} \times \frac{0.105 \text{ mol NaOH}}{1000 \text{ mL}} \times \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}}}{60 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.0212625 \text{ M}$$

9. For the reaction, $\text{HCO}_2\text{H}(\text{aq}) + \text{CN}^-(\text{aq}) \rightleftharpoons \text{HCO}_2^-(\text{aq}) + \text{HCN}(\text{aq})$, what is the equilibrium constant and does it favor the formation of reactants or products? The acid dissociation constant, K_a , for HCO_2H is 1.8×10^{-4} and the acid dissociation constant for HCN is 4.0×10^{-10} .

D

- A. $K = 1.00$. The reaction favors neither the formation of reactants nor products.
 B. $K = 2.2 \times 10^{-6}$. The reaction favors the formation of products.
 C. $K = 2.2 \times 10^{-6}$. The reaction favors the formation of reactants.
 D. $K = 4.5 \times 10^5$. The reaction favors the formation of products.
 E. $K = 4.5 \times 10^5$. The reaction favors the formation of reactants.

$$K_{eq} = 10^{\Delta pK_a}$$

$$\Delta pK_a = pK_a(\text{acid product}) - pK_a(\text{acid reactant})$$

$$= (-\log(4.0 \times 10^{-10})) - (-\log(1.8 \times 10^{-4}))$$

$$= 5.653212514$$

$$\therefore K_{eq} = 10^{5.653212514} = 4.500000002 \times 10^5$$

↓
product favored
(weak acid & weak base are favored)

10. What happens to a $\text{HC}_2\text{H}_3\text{O}_2/\text{C}_2\text{H}_3\text{O}_2^-$ buffer if we add $\text{HC}_2\text{H}_3\text{O}_2$ to it?

B

- A. The pH will go up.
 B. The pH will go down.
 C. The pH will not change.



weak acid \therefore
pH will go down

Formula Sheet

Length

1 kilometer = 0.62137 mile
1 inch = 2.54 centimeters (exactly)
1 Ångstrom = 1×10^{-10} meter

Energy

1 joule = $1 \text{ kg} \cdot \text{m}^2 / \text{s}^2$
1 calorie = 4.184 joules
1 Calorie = 1 kilocalorie = 1000 calories
1 L·atm = 101.325 joules

Pressure

1 pascal = $1 \text{ N} / \text{m}^2 = 1 \text{ kg} / \text{m} \cdot \text{s}^2$
1 atmosphere = 101.325 kilopascals = 760 mm Hg = 760 torr = 14.70 lb/in²
1 bar = 1×10^5 Pa (exactly)

Temperature

0 K = -273.15°C
K = °C + 273.15
°C = (5/9)(°F - 32)

Mass

1 kg = 2.205 lbs

Volume

1 mL = $1 \text{ cm}^3 = 1 \text{ cc}$

Constants

$c = 2.998 \times 10^8 \text{ m/sec}$
 $h = 6.626 \times 10^{-34} \text{ J} \cdot \text{sec}^{-1}$
 $R = 0.08206 \text{ L} \cdot \text{atm} / \text{mol} \cdot \text{K} = 8.314 \text{ J} / \text{mol} \cdot \text{K}$
Specific heat of water = 4.184 J/g·K
Mass of an electron: $9.109 \times 10^{-31} \text{ kg}$
Mass of a proton: $1.673 \times 10^{-27} \text{ kg}$
 $RH = 2.18 \times 10^{-18} \text{ J}$
Specific heat of water = 4.184 J/g·K
STP = 273.15 K and 1 atm
Avogadro's number: 6.022×10^{23}

Equations

$d \text{ (density)} = m/V$
 $P_1 V_1 = P_2 V_2$
 $V_1/T_1 = V_2/T_2$
 $P_1 V_1/n_1 T_1 = P_2 V_2/n_2 T_2$
 $PV = nRT$
 $(P + a(n^2/V^2)) \cdot (V - nb) = nRT$
molar mass (M) = mRT/PV
density (d) = MP/RT
 $x_A = n_A/n_{\text{tot}} = P_A/P_{\text{tot}} = V_A/V_{\text{tot}}$
 $P_{\text{tot}} = P_A + P_B + \dots$
 $n_{\text{tot}} = n_A + n_B + \dots$
$$\mu_{rms} = \sqrt{\frac{3RT}{M}}$$

$$\frac{\text{Rate of effusion A}}{\text{Rate of effusion B}} = \sqrt{\frac{MW_B}{MW_A}}$$

$$Q = C \times \Delta T = c_{\text{specific}} \times m \times \Delta T$$

$$Q = n \times \Delta H \text{ (kJ/mol)} = m \times \Delta H \text{ (kJ/g)}$$

$$w = -P\Delta V$$

$$\Delta E = q + w$$

$$\Delta H^\circ = \sum n\Delta H_f^\circ(\text{products}) - \sum n\Delta H_f^\circ(\text{reactants})$$

$$\Delta H^\circ = \sum n\Delta H^\circ(\text{bonds broken}) - \sum n\Delta H^\circ(\text{bonds formed})$$

$$E = h\nu$$

$$c = \lambda\nu$$

$$\lambda = h/mv$$

$$\Delta E = -2.18 \times 10^{-18} J \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

$$\ln\left(\frac{P_2}{P_1}\right) = \frac{\Delta H_{vap}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

$$C_g = kP_g$$

$$P_{\text{solution}} = P_{\text{solvent}} X_{\text{solvent}}$$

$$P_{\text{solution}} = \sum P_j = \sum P_j X_j$$

$$\Delta T_b = K_b m_i$$

$$\Delta T_f = K_f m_i$$

$$\pi = MRTi$$

Thermodynamic and Electrochemistry

$$S = k_b \times \ln(W)$$

$$k_b = 1.381 \times 10^{-23} \text{ J/K}$$

$$\Delta S = q_{\text{rev}}/T$$

$$\Delta S_{\text{surr}} = q_{\text{surr}}/T = -q_{\text{rev}}/T$$

$$\Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}}$$

$$\Delta S^\circ_{\text{rxn}} = \sum \nu S^\circ_{\text{products}} - \sum \nu S^\circ_{\text{reactants}}$$

$$\Delta H^\circ_{\text{rxn}} = \sum \nu H^\circ_{\text{products}} - \sum \nu H^\circ_{\text{reactants}}$$

$$\Delta G^\circ_{\text{rxn}} = \sum \nu G^\circ_{\text{products}} - \sum \nu G^\circ_{\text{reactants}}$$

$$\Delta G = \Delta H - T\Delta S$$

$$\Delta G = \Delta G^\circ + RT \cdot \ln Q$$

$$R = 8.314 \text{ J/mol.K}$$

$$\Delta G^\circ = -RT \cdot \ln K$$

$$\Delta G = -nFE_{\text{cell}}$$

$$F = 96485 \text{ J/(V}\cdot\text{mol e}^-)$$

$$E^\circ_{\text{cell}} = RT/nF \ln K$$

$$E^\circ_{\text{cell}} = (0.0257/n) \ln K = (0.0592/n) \log K$$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - (RT/nF) \ln Q$$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - (0.0257/n) \ln Q$$

$$\text{Electrolysis: } Q \text{ (total charge)} = I \times t = n \times F$$

Integrated Rate Laws & half-life

$$\ln \frac{[A]}{[A]_0} = -kt$$

$$\frac{1}{[A]} = kt + \frac{1}{[A]_0}$$

$$[A] = -kt + [A]_0$$

$$t_{1/2} = \frac{[A]_0}{2k}$$

$$t_{1/2} = \frac{\ln 2}{k} = \frac{0.693}{k}$$

$$t_{1/2} = \frac{1}{k[A]_0}$$

$$\ln \frac{k_2}{k_1} = -\frac{E_a}{R} \left(\frac{1}{T_2} - \frac{1}{T_1} \right)$$

Equilibrium and Acid / Base

$$K_p = K_c \times (RT)^{\Delta n}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$K_w = 1.0 \times 10^{-14} \text{ at } 25^\circ\text{C}$$

$$K_w = [\text{H}_3\text{O}^+] \times [\text{OH}^-]$$

$$K_w = K_a \times K_b$$

$$\text{p}K_a = -\log[K_a]$$

$$\text{Buffer: pH} = \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

$$\ln \frac{K_2}{K_1} = \frac{\Delta H_{rxn}^\circ}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

Periodic Table of the Elements

1																		18																																																											
1 H 1.01																		2 He 4.00																																																											
2																		13																																																											
3 Li 6.94				4 Be 9.01				5										6				7				8				9				10				11				12				13				14				15				16				17				18											
11 Na 22.99				12 Mg 24.31				3										4				5				6				7				8				9				10				11				12				13				14				15				16				17				18			
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 65.38				31 Ga 69.72				32 Ge 72.63				33 As 74.92				34 Se 78.97				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 95.95				43 Tc [97]				44 Ru 101.07				45 Rh 102.91				46 Pd 106.42				47 Ag 107.87				48 Cd 112.41				49 In 114.82				50 Sn 118.71				51 Sb 121.76				52 Te 127.60				53 I 126.90				54 Xe 131.29									
37 Cs 132.91				56 Ba 137.33				72 Hf 178.49										73 Ta 180.95				74 W 183.84				75 Re 186.21				76 Os 190.23				77 Ir 192.22				78 Pt 195.08				79 Au 196.97				80 Hg 200.59				81 Tl 204.38				82 Pb 207.2				83 Bi 208.98				84 Po [209]				85 At [210]				86 Rn [222]							
87 Fr [223]				88 Ra [226]				104 Rf [267]										105 Db [268]				106 Sg [269]				107 Bh [270]				108 Hs [269]				109 Mt [277]				110 Ds [281]				111 Rg [282]				112 Cn [285]				113 Nh [286]				114 Fl [290]				115 Mc [290]				116 Lv [293]				117 Ts [294]				118 Og [294]							

57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm [145]	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97
89 Ac [227]	90 Th 232.04	91 Pa 231.04	92 U 238.03	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]	103 Lr [262]