The correct answers are highlighted in blue.
The answer key is at the end.

1. The **pH** of a 0.01 M aqueous solution of a strong base is:

   A) 3
   B) 13
   C) 12
   D) 1
   E) 7

With a strong base, $[HO^-]$ = [base].

$[HO^-]$ = 0.01 M.

pOH = -log(0.01)
      = 2

pH   = 14 - pOH
       = 12

---

2. **Which** of the following combinations of aqueous solutions would give a **buffer**? (All
   solutions are 1.0 M).

   A)   10 mL $HClO_4$ + 20 mL $KClO_2$
   B)   15 mL $HClO_4$ + 15 mL $KClO_4$
   C)   20 mL $HClO_4$ + 10 mL KOH
   D)   20 mL $HClO_2$ + 10 mL $KClO_4$
   E)   10 mL $HClO_2$ + 20 mL KOH

A)   Limiting strong acid ($HClO_4$) plus weak base ($KClO_2$) gives a buffer.

B)   Strong acid ($HClO_4$) plus neutral salt ($KClO_4$) gives a strong acid solution.

C)   Strong acid ($HClO_4$) plus limiting strong base (KOH) gives a strong acid
     solution.

D)   Weak acid ($HClO_2$) plus neutral salt ($KClO_4$) gives a weak acid solution.

E)   Weak acid ($HClO_2$) plus excess strong base (KOH) gives a solution of a weak
     base salt (which has a negligible effect on pH) plus a strong base solution.
     Overall, it is a strong base solution.

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3. Which of the following compounds will produce a **basic solution** when placed into water?

(i) Na_2O (ii) NH_4Cl (iii) LiBr (iv) CH_3COOK (v) CH_3NH_2

A) iii, v
B) i, ii
C) ii, iii, iv
D) ii, iv
E) i, iv, v

(i) Strong base (O^{2-})

(ii) Weak acid (NH_4^+)

(iii) Neutral salt (Li^+ and Br^-)

(iv) Weak base salt (CH_3COO^-)

(v) Weak base (an amine, $R-NH_2$)

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4. MES is a common buffer in chemical biology. At pH 6.5, the ratio of MES, a weak base, to its conjugate acid, $MESH^+$, is $[MES]/[MESH^+] = 2.19$. What is the **pK_a** of $MESH^+$?

A) 7.17
B) 7.47
C) 6.76
D) 5.65
E) 6.16

Henderson-Hasselbalch equation: $pH = pK_a + \log([A^-]/[HA])$

Rearrange:
$pK_a = pH - \log([A^-]/[HA])$
$= 6.5 - \log(2.19)$
$= 6.16$

5. Which of the following statements about buffers is **false**?

A) The percent ionization of the weak acid or weak base in a buffer is negligible (close to zero).
B) Weak acid-strong base titration curves show a buffer region right up until the equivalence point.
C) Buffers resist changes in pH by converting strong base into weak base, or strong acid into weak acid.
D) The buffer capacity for a weak acid (HA) / weak base (A^-) system is at a maximum when [HA] and [A^-] are equal.
E) A buffer's capacity can be exceeded by adding excess amounts of strong acid.

(A) True. This is already true for a weak acid or a weak base alone, but in a buffer, the common ion effect further suppresses ionization.

(B) False. Our definition of a buffer includes the requirement that 0.1 < [A^-]/[HA] < 10. As a weak acid/strong base titration nears the equivalence point, there is a region before the equivalence point where [A^-]/[HA] > 10. There is too little remaining HA for the solution to be an effective buffer.

(C) True. For example, when a limiting amount of NaOH is added to a benzoic acid/benzoate buffer, HO^- reacts with benzoic acid to produce an equal amount of benzoate ion.

(D) True. The buffer capacity is maximal when the concentrations of both A^- and HA are large and equal.

(E) True. Once one equivalent of strong acid has been added, the solution is now a weak acid solution, not a buffer. Any further addition of strong acid makes it a strong acid solution.

6. What is $[H_3O^+]$ **(in M)** in 1.0 M $CH_3COOH(aq)$?

Data:
$K_a(CH_3COOH) = 1.8 \times 10^{-5}$

A) 1.66×10^{-3}
B) 1.34×10^{-3}
C) 3.22×10^{-3}
D) 4.24×10^{-3}
E) 7.18×10^{-3}

This is a weak acid problem.

$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$

	CH_3COOH	H_2O	CH_3COO^-	H_3O^+
I	1.0 M	-	0 M	0 M
C	- x M	-	+ x M	+ x M
E	(1.0 - x) M	-	x M	x M

Make the small x assumption (i.e., x << 1.0)

$K_a = [A^-][H_3O^+]/[HA] = x^2/[HA]$.

Therefore:
$x = \sqrt{\{K_a \times [HA]\}}$
$= \sqrt{\{1.8 \times 10^{-5} * 1.0\}}$
$= 4.24 \times 10^{-3}$ M

$[H_3O^+] = 4.24 \times 10^{-3}$ M

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7. A buffer is made using only Tris (weak base) and TrisHCl (weak acid). **What mass (in g)** of solid TrisHCl would be used to make 0.20 L of a buffer solution with a total buffer concentration ([Tris] + [TrisH$^+$]) of 0.050 M, and pH 7.80?

$TrisH^+ + H_2O \rightleftharpoons Tris + H_3O^+$
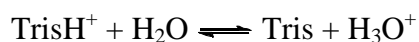
Data:
$pK_a(TrisH^+) = 8.30$
mol wt(TrisHCl) = 157.6 g/mol
mol wt(Tris) = 121.1 g/mol

A) 10.5
B) 2.80
C) 1.20
D) 1.50
E) 0.80

First, calculate the required concentration of TrisHCl (HA) from the $[A^-]/[HA]$ ratio (Henderson-Hasselbalch equation) and the total required concentration of ($[A^-]$+ [HA]):

$[A^-]/[HA]$ $= 10^{(pH-pKa)}$
                 $= 10^{(7.8-8.3)}$
                 $= 0.316$

Therefore:
$[A^-] = 0.316[HA]$

and

$[A^-] + [HA] = 0.05$ M

so

$1.316[HA] = 0.05$ M

and

$[HA] = 0.038$ M

We require a 0.2 L solution, so:
n(TrisHCl) $= 0.038$ mol/L $* 0.2$ L
                $= 0.0076$ mol

mass(TrisHCl) $= 0.0076$ mol $* 157.25$ g/mol
                $= 1.20$ g

8. What is the **pH** of the solution that results when the following substances are added to water to create 1.00 L of solution?

0.100 mol HNO_3
0.100 mol HNO_2
0.300 mol $NaNO_2$

Data:
$K_a(HNO_2) = 7.2 \times 10^{-4}$

A) 3.53
B) 4.65
C) 6.04
D) 2.30
E) 3.14

If we consider that the strong acid (HNO_3) will react with the weak base ($NaNO_2$) then we can construct the following mol table:

	HNO_3	$NaNO_2$	→	HNO_2	$NaNO_3$
before	0.100	0.300		0.100	0
after	0	0.200		0.200	0.100

HNO_3 is completely consumed, forming $NaNO_3$, a neutral salt. The remaining solution is a buffer.

Use the Henderson-Hasselbalch equation:

pK_a = -log(7.2×10^{-4}) = 3.14

pH = pK_a + log([A⁻]/[HA])
= 3.14 + log(0.200/0.200)
= 3.14

---

9. Bromophenol blue is an acid-base indicator with a pK_a of 4.0. The acidic form is yellow and the basic form is blue-violet. A few drops of this indicator are added to the titration of benzoic acid (beaker, $pK_a = 4.20$) with $NaOH(aq)$ (buret). Find the **false** statement(s) about this titration experiment.

(i)   The indicator starts to change from yellow to blue-violet at about pH 3 and completes the change from to blue-violet at about pH 5.
(ii)   The indicator will be yellow over the entire buffer region of the titration curve.
(iii)   The bromophenol blue indicator is not an appropriate choice for locating the equivalence point of this titration.

A) No statements are false
B) i
C) ii
D) iii
E) i, iii

(i)   True. A pH indicator changes colour over the range of pH = pK_{HIn} - 1 to pH = pK_{HIn} + 1.

(ii) False. Benzoic acid is a buffer in the region pH = pK_a ± 1, or 3.2 to 5.2. At pH 5.2, bromophenol blue will be blue-violet.

(iii) True. An appropriate indicator would have pK_{HIn} close to the pH at the equivalence point, which would be > 7 in a weak acid/strong base titration.

10. Which would be the **best pH indicator** to find the endpoint in a titration of 0.050 M cacodylic acid with 0.100 M LiOH? (Cacodylic acid is a monoprotic acid.)

Data:
K_a(cacodylic acid) $= 5.7 \times 10^{-7}$

Indicator	Colour change range (pH)
A) Methyl orange	3.1 - 4.4
B) Thymolphthalein	9.3 - 10.5
C) Cresol red	7.2 - 8.8
D) Bromophenol blue	3.0 - 4.6
E) Bromothymol blue	6.0 - 7.6

We need to know the pH at the equivalence point before deciding on an appropriate pH indicator.

In a weak acid/strong base titration, the product at the equivalence point will be a weak base salt. Given the relative concentrations of cacodylic acid and LiOH, it will take 1/2 of a volume of LiOH to neutralize the weak acid. The final volume will be 1.5 times the starting volume, and the lithium cacodylate concentration will be 0.05 M ÷ 1.5 = 0.033 M.

We need the K_b value to solve the equation for a weak base:

K_b(lithium cacodylate) $= K_w/K_a$(cacodylic acid)
$\qquad\qquad\qquad\qquad\quad = 10^{-14} / 5.7 \times 10^{-7}$
$\qquad\qquad\qquad\qquad\quad = 1.75 \times 10^{-8}$

We now set up an ICE table, with A^- = cacodylate$^-$ and HA = cacodylic acid:

	A^- (aq) +	H_2O(l) $\rightleftharpoons$	HA(aq)	+ OH^-(aq)
I	0.033 M	-	0 M	0 M
C	- x M	-	+ x M	+ x M
E	(0.033 - x) M	-	x M	x M

Make the small x assumption (i.e., x << 0.033):

$K_b = [HA][HO^-]/[A^-] = x^2/[A^-]$.

Therefore:
x $= \sqrt{\{K_b \times [A^-]\}}$
$\quad = \sqrt{\{1.75 \times 10^{-8} * 0.033\}}$
$\quad = 2.41 \times 10^{-5}$ M
$\quad = [HO^-]$

pH $= 14 - pOH$
$\quad = 14 - (-\log(2.41 \times 10^{-5}))$
$\quad = 9.38$

Thymolphthalein is the only indicator whose colour change occurs at close to the equivalence point.

11. Some parts of human cells, like the endosomes, are acidic, with pH ≈ 5.5. **Which** of the following **weak acids** would be the most effective acid component of a buffer to maintain pH 5.5 (assuming equal concentrations of all buffers)?

	Weak acid	K_a
A)	glycine	4.6×10^{-3}
B)	lactic acid	1.4×10^{-4}
C)	$H_2PO_4^-$	6.3×10^{-8}
D)	H_2CO_3	4.5×10^{-7}
E)	NH_4^+	5.5×10^{-10}

The best buffer will be the one with pK_a closest to 5.5.

Weak acid pK_a
glycine   2.34
lactic acid 3.85
$H_2PO_4^-$ 7.20
H_2CO_3 6.35
NH_4^+ 9.26

Only carbonic acid has a pKa within 1 pH unit of the target pH.

12. If the pH indicator cresol red (yellow at low pH, purple at high pH) is placed in a buffer solution with pH 7.40, **what percentage** of the indicator would be in the protonated (HIn) form?

Data:
pK_{HIn}(cresol red) = 8.32

A) 19%
B) 79%
C) 9%
D) 39%
E) 89%

This is can be answered like a buffer question because pH indicators are weak acids/bases.

Rearrange the Henderson-Hasselbalch equation (pH = pK_a + log([A⁻]/[HA])):

[A⁻]/[HA]   = $10^{(pH-pKa)}$
              = $10^{(7.4-8.32)}$
              = 0.12
or
              [A⁻] = 0.12[HA]

We want to know the percent of the HA form:

```
% HA    = [HA]/([A⁻] + [HA]) * 100%
        = [HA]/(0.12[HA] + [HA]) * 100%
        = 1/1.12 * 100%
        = 89%
```

13. Which of the following titration graphs is **completely accurate**?

Data:
K_a(benzoic acid) $= 6.3 \times 10^{-5}$
K_b(ammonia) $= 1.8 \times 10^{-5}$

(A) titrate benzoic acid with NaOH

(B) titrate benzoic acid with NaOH

(C) titrate benzoic acid with HCl

(D) titrate sodium benzoate with HCl

(E) titrate ammonia with HCl
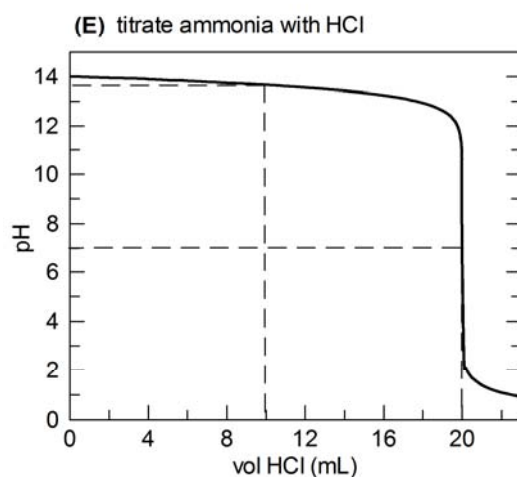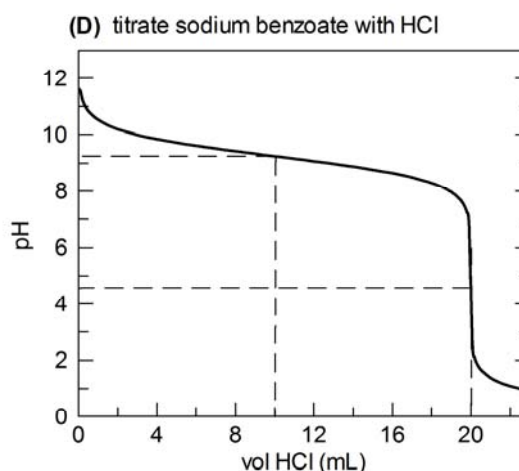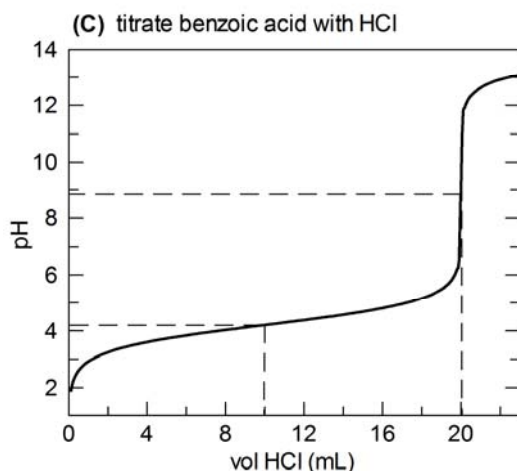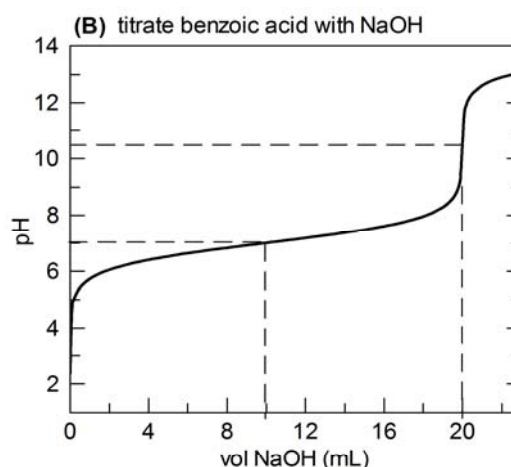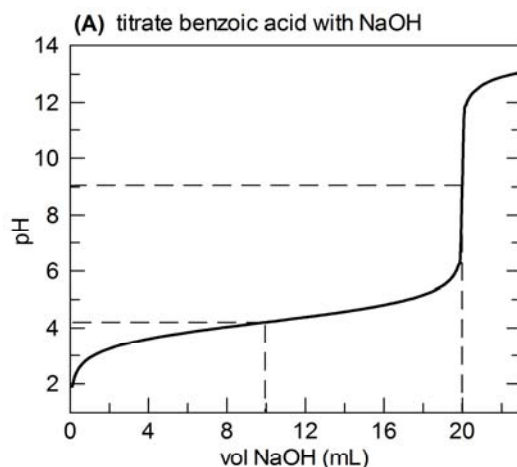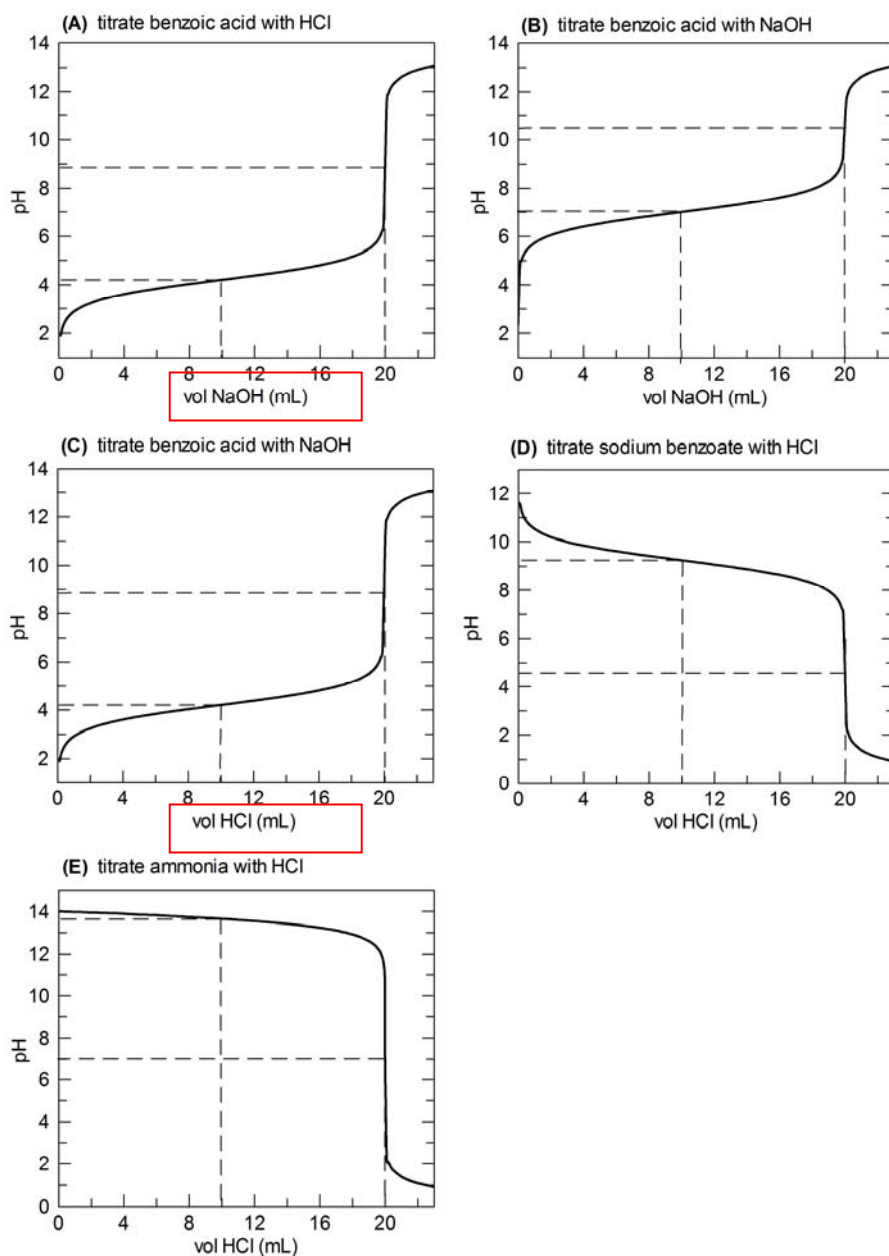

(A) Correct.

(B) Incorrect. The 1/2-equivalence point should be at the pK_a (4.2), not pH 7.

(C) Incorrect. Titrating a weak acid with a strong acid.

(D) Incorrect. At the 1/2-equivalence point pH = pK_a = 4.2, not pH 9.

(E) Incorrect. The graph shows a strong base/strong acid titration with no buffer region. The 1/2-equivalence point for an NH$_3$/HCl titration would be at pK_a(NH$_4^+$) = 9.26.

In versions 2 and 4, the x-axes (abscissa) in parts A and C were incorrectly labelled, as announced during the test. The titles were correct. In part A, the abscissa should have been "vol HCl (mL)", and in part C, it should have been "vol NaOH (mL)". It is now apparent that this announcement was not clear in all test rooms, and so answer A will also be awarded full marks in order to account for those who changed the question titles to match the titration curve axes. Students in the early/alternative write rooms did not receive any announcement, thus they will receive full marks for any answer because of the time investment in the question and lack of correct answer in the absence of the announcement.

(A) titrate benzoic acid with HCl

(B) titrate benzoic acid with NaOH

(C) titrate benzoic acid with NaOH

(D) titrate sodium benzoate with HCl

(E) titrate ammonia with HCl

12

14. The titration of 10.00 mL of a weak base with a 0.180 M HCl(aq) solution reaches the half-equivalence point after adding 4.75 mL of acid. The pH at the half-equivalence point is 9.42. Calculate the **initial concentration of the base** and its K_b value.

A) 0.187 M $K_b = 2.6 \times 10^{-10}$
B) 0.171 M $K_b = 2.6 \times 10^{-5}$
C) 0.187 M $K_b = 2.6 \times 10^{-5}$
D) 0.171 M $K_b = 3.8 \times 10^{-10}$
E) 0.200 M $K_b = 3.8 \times 10^{-10}$

---

n(HCl) required to reach the 1/2-equivalence point:

n(HCl) = 0.180 mol/L * 4.75 × 10^{-3} L (or 0.180 mol/L * 4.75 mL * L/1000 mL)
       = 8.55 × 10^{-4} mol

At the 1/2-equivalence point, [A⁻] = [HA], so the original n(base) is double that:

n(base)   = 2 × 8.55 × 10^{-4} mol
                = 1.71 × 10^{-3} mol

[base]   = 1.71 × 10^{-3} mol / 0.01000 L
                = 0.171 M

At the 1/2-equivalence point, pH = pK_a = 9.42, therefore:

pK_b = 14 - pK_a
     = 4.58

K_b = 10^{-pK_b}
      = $10^{-4.58}$
      = 2.6 × 10^{-5}

---

15. In a titration of 0.050 M NaOH (beaker) with 0.20 M HCl (buret), what is the **pH** at the half-equivalence point?

A) 2.53
B) 3.33
C) 7.00
D) 11.55
E) 12.35

We are not given the volumes of each solution, so we will solve this using relative volumes. If the original volume of NaOH is v_1, reaching the 1/2-equivalence point, it would require:

v_{added} = $v_1(c_1/c_2)$ * 0.5 (for 1/2-equiv.)
           = 0.050 M / 0.20 M * 0.5 * v_1
           = 0.125v_1

Therefore:

v_2 = v_1 + v_{added}
= $1.125v_1$

At the 1/2-equivalence point:

[NaOH] = $0.5 \times c_1 \times v_1/v_2$
= 0.5×0.05 M $\times 1/1.125$
= 0.022 M
= [HO⁻]

pOH = -log(0.022)
= 1.65

pH = 14 - pOH
= 12.35

16. A 1.00 L solution of a HCOOH/HCOONa buffer has pH 3.60, and [HCOONa] = 0.68 M. After 50.0 mL of 1.00 M HCl are added, what is the **new pH**?

Data:
K_a(HCOOH) = 1.8×10^{-4}

A) 3.49
B) 3.39
C) 3.44
D) 3.54
E) 3.61

pK_a(HCOOH) = -log(1.8 × 10⁻⁴)
= 3.74

Rearrange the Henderson-Hasselbalch equation as in question 12:

[A⁻]/[HA] = $10^{(pH-pKa)}$

or

[HA] = [A⁻]/$10^{(pH-pKa)}$
= 0.68 M/$10^{(3.60 - 3.74)}$
= 0.95 M

n(HCl) = 0.0500 L * 1.00 mol/L
= 0.0500 mol

Create a mol table to see how much weak base is converted to weak acid. We started with a 1.00 L solution, so n(HCOOH) = [HCOOH], and n(HCOONa) = [HCOONa].

```
         HCOONa    +    HCl    →     HCOOH    +    NaCl
before   0.68           0.0500       0.95           0
after    0.63           0            1.00           0.05
```

$pH = pK_a + \log([HCOONa]/[HCOOH])$
$= 3.74 + \log(0.63/1.00)$
$= 3.54$

17. A student titrates a weak acid (beaker, 20. mL, 0.22 M, $K_a = 1.0 \times 10^{-4}$) with NaOH (buret, 0.12 M). Which of the following statements would be **incorrect** regarding this titration?

    A) The initial pH of the solution in the beaker would be below 4.
    B) The equivalence point should occur at 37 mL of added NaOH.
    C) The buffer region would occur at a pH above 7.
    D) The half equivalence point would occur at pH 4.
    E) At the equivalence point, the pH of the solution would be basic.

(A) Correct. This is a weak acid solution, pH = 2.33. (Solve as for question 6.)

(B) Correct. n(weak acid) = 0.02 L * 0.22 mol/L = 0.0044 mol
            vol(NaOH) = 0.0044 mol / 0.12 mol/L = 0.037 L, or 37 mL.

(C) Incorrect. $pK_a = -\log(1.0 \times 10^{-4}) = 4.0$
            Therefore, the buffer region will be pH 3.0 to 5.0.

(D) Correct. At the 1/2-equivalence point, pH = pK_a = 4.0 (see part C).

(E) Correct. At the equivalence point, the weak acid would have been converted
             to a weak basic salt solution, with pH > 7.

18. Which statement is **incorrect** regarding the pseudo-zero order reaction A → G?

    A) A graph of v_0 vs. [A] is a horizontal line.
    B) $v = k[A]$
    C) average rate = $-\Delta[A]/\Delta t$
    D) A graph of [A] vs. time is linear.
    E) The plot of v vs. time is a horizontal line.

(A) Correct. v_0 does not vary with [A], so the plot will be a horizontal line.

(B) Incorrect. $v = k$; it does not depend on [A].

(C) Correct. The negative sign is used because we're following the decrease in
            reactant, A.

(D) Correct. The rate does not vary with [A], so [A] decreases at a constant

(E) Correct. v does not change with time, so the plot is horizontal.

19. The table below shows the rate of formation of G during the reaction A → 2G. What is the **rate constant** for this reaction?

time (h)	[G] (M)
1.0	0.12
2.0	0.24
4.0	0.48

A) $3.0 \times 10^{-2} \text{ h}^{-1}$
B) $6.0 \times 10^{-2} \text{ M h}^{-1}$
C) 3.0 M h^{-1}
D) 0.030 M h^{-1}
E) 0.060 h^{-1}

[G] increases linearly with time (always 0.12 M/h), so this is a zero order reaction. For a zero order reaction $v = k$.

$k = v = 1/g \times \Delta G/\Delta t$
$= \frac{1}{2} \times 0.12 \text{ M} / 1.0 \text{ h}$
$= 0.06 \text{ M/h or } 6.0 \times 10^{-2} \text{ M h}^{-1}$.

20. Which statement regarding the rate of the reaction 2A + B → G + H is **incorrect**?

A)   The reaction rate can be described as: average rate = $\Delta[H]/\Delta t$.
B)   The reaction rate can be described as: average rate = $-(1/2 \times \Delta[A])/\Delta t$.
C)   The instantaneous reaction rate can be described as: $v = -d[B]/dt$.
D)   If the reaction rate doubles with a doubling of either [A] or [B], then the rate law may written as $v = k[A][B]$.
E)   The reaction order can be determined from the reaction stoichiometry shown above.

(A) Correct.   The symbol Δ is used for average rates to denote finite differences in time.

(B) Correct.   It is negative because we are following the decrease in reactant concentration, and 1/2 because the stoichiometry is 2 equivalents of A per equivalent of B.

(C) Correct.   "d" is used to indicate the limit as $\Delta t \rightarrow 0$.

(D) Correct.   The reaction is first order with respect to (w.r.t.) both A and B, and second order overall.

(E) Incorrect. There is not (necessarily) any relationship between stoichiometry

and the rate law.

21. Based on the observed rates, what is the **rate law** for the reaction $A + B \rightarrow G$?

[A] (M)	[B] (M)	v_0 (M/s)
0.1	0.1	0.4
0.1	0.3	1.2
0.2	0.2	3.2

A) $v_0 = k[A][B]$

B) $v_0 = k[A]^2[B]$

C) $v_0 = k[A][B]^2$

D) $v_0 = k[A]^0[B]$

E) $v_0 = k[A]^0[B]^2$

Comparing lines 1 and 2, the rate triples when [B] triples, so the reaction is first order with respect to (w.r.t.) [B].

Given the fact that the reaction is first order w.r.t. [B], the reaction rate would be 0.8 M/s in line 3 if the reaction were zero order w.r.t. [A]. However, doubling [A] further quadrupled v_0 to 3.2 M/s, meaning that the reaction is second order w.r.t. [A].

More formally:

v_0(line3)/v_0(line 1) $= k[A_3]^m[B_3]^1/ k[A_1]^m[B_1]^1$ (cancel k's)

3.2 M/s / 0.4 M/s $= \{(0.2\ M)^m \times (0.2\ M)^1\} / \{(0.1\ M)^m \times (0.1\ M)^1\}$

8                              $= 2^m \times 2$ (the A and B terms must be treated separately because if m and n are not equal, then the units (M^2 and M) will not be equal)

therefore:

m = 2

22. Which statement regarding reaction rates is **incorrect**?

    A) Initial rates can be approximated by measuring average rates at closely spaced time points, and as close to time t = 0 as possible.
    B) Instantaneous rates at time t can be approximated by measuring average rates over a very small time interval about t.
    C) Experimentally determined rates are always average rates.
    D) Two reactions that have identical average rates over one time period must have identical average rates over all subsequent time periods.
    E) An instantaneous rate describes the rate of a reaction at a specific time, t.

(A) Correct.

(B) Correct.

(C) Correct.

(D) Incorrect.   As seen in the notes, many reactions can have an average rate of 1 M/min over the first minute, but different rates over other time periods, depending on the order of the reaction.

(E) Correct.

23. The rate law for the reaction $A + 2B \rightarrow 2C$ was determined to be $v = k[A][B]^2$, with a rate constant $k = 1.2 \times 10^{-2} \, M^{-2}s^{-1}$ at 20°C. What is the **reaction rate (in Ms^{-1})** when $[A] = 0.1$ M and $[B] = 0.2$ M?

A) 2.6×10^{-4}
B) 4.4×10^{-4}
C) 4.8×10^{-5}
D) 1.2×10^{-5}
E) 2.2×10^{-5}

v_0 $= k[A][B]^2$
     $= (1.2 \times 10^{-2} \, M^{-2}s^{-1}) \, * \, (0.1 \, M) \, * \, (0.2 \, M)^2$
     $= 4.8 \times 10^{-5} \, Ms^{-1}$

24. Which of the following reactions is **not third order** overall?

A) $v_0 = k[B]^2[C]^0[D]$

B) $v_0 = k[B][C][D]$

C) $v_0 = k[B]^3[C]^0[D]^0$

D) $v_0 = k[A]^2[C][D]$

E) $v_0 = k[A][C]^2[D]^0$

(D) Is 4th order overall: $v_0 = k[A]^2[C][D]$

Answer key, version 1.

1. C
2. A
3. E
4. E
5. B
6. D
7. C
8. E
9. C
10. B
11. D
12. E
13. A
14. B
15. E
16. D
17. C
18. B
19. B
20. E
21. B
22. D
23. C
24. D

Answer key, version 2.

1. A
2. D
3. A
4. A
5. E
6. E
7. C (full marks will be awarded for answer A also)
8. B
9. A
10. B
11. D
12. E
13. D
14. B
15. A
16. D
17. A
18. E
19. D
20. A
21. E
22. B
23. A
24. C

Answer key, version 3.

1. D
2. D
3. A
4. B
5. B
6. E
7. A
8. A
9. C
10. D
11. C
12. A
13. C
14. B
15. E
16. B
17. C
18. B
19. B
20. D
21. B
22. A
23. B
24. E

Answer key, version 4.

1. E
2. B
3. A
4. C
5. C
6. E
7. C (full marks will be awarded for answer A also)
8. E
9. B
10. A
11. E
12. B
13. E
14. A
15. C
16. E
17. B
18. E
19. D
20. D
21. E
22. A
23. C
24. B