

The answer keys appear at the end of the document.

VERSION 1.

Section #1 – These questions are worth two marks each.

1. The **pH** of a 0.1 M solution of a strong acid is:

- A) 12
- B) 13
- C) 2
- D) 7
- E) 1

Strong acids are completely ionized:



Therefore $[\text{H}_3\text{O}^+] = [\text{HA}] = 0.1 \text{ M}$.

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(0.1) = 1$$

2. The pH of a $\text{PO}_4^{3-}/\text{HPO}_4^{2-}$ system is 12.12. What is the **ratio of $[\text{PO}_4^{3-}]/[\text{HPO}_4^{2-}]$** in this solution?

Data: $K_a(\text{HPO}_4^{2-}) = 4.2 \times 10^{-13}$

- A) 0.98
- B) 1.9
- C) 0.55
- D) 2.4
- E) 0.41

Henderson-Hasselbalch equation:

$$\text{pH} = \text{p}K_a + \log([\text{A}^-]/[\text{HA}])$$

therefore:

$$[\text{A}^-]/[\text{HA}] = 10^{(\text{pH} - \text{p}K_a)}$$

$$\text{pH} = 12.12$$

$$\begin{aligned}\text{p}K_a &= -\log(4.2 \times 10^{-13}) \\ &= 12.38\end{aligned}$$

therefore:

$$\begin{aligned}[\text{A}^-]/[\text{HA}] &= 10^{(12.12 - 12.38)} \\ &= 0.55\end{aligned}$$

3. Which of the following combinations of solutions can be used to **prepare a buffer solution**? (All solutions are 1.0 M)

- A) 25 mL HCl + 25 mL KNO₂
- B) 25 mL HNO₃ + 15 mL NaOH
- C) 15 mL NaOH + 25 mL HNO₂
- D) 25 mL NaOH + 25 mL KNO₂
- E) 15 mL HCl + 25 mL KNO₃

- A) 25 mL HCl + 25 mL KNO₂ - equimolar strong acid + weak base - not a buffer
- B) 25 mL HNO₃ + 15 mL NaOH - strong acid + strong base - not a buffer
- C) 15 mL NaOH + 25 mL HNO₂ - limiting amount of strong base + weak acid - **buffer**
- D) 25 mL NaOH + 25 mL KNO₂ - strong base + weak base - not a buffer
- E) 15 mL HCl + 25 mL KNO₃ - strong acid + conjugate base of a strong acid (very weak base) - not a buffer

4. The **effect on pH** of adding 20.0 mL of 0.010 M NaOH(aq) to 1.00L of a solution that is 0.20 M NH₃(aq) and 0.20 M NH₄Cl(aq) is to:

- A) lower the pH by several units
- B) have no effect on pH
- C) lower the pH slightly
- D) **raise the pH slightly**
- E) raise the pH by several units

The solution is a buffer before addition of the strong base because (i) conjugate weak acid/weak base pair, (ii) $0.1 < [A^-]/[HA] < 10$, and (iii) $[HA]/K_a > 100$.

NH₄Cl: $0.20 \text{ M} \times 1.00 \text{ L} = 0.2 \text{ mol}$

NH₃: $0.20 \text{ M} \times 1.00 \text{ L} = 0.2 \text{ mol}$

NaOH: $0.010 \text{ M} \times 0.0200 \text{ L} = 2.0 \times 10^{-4} \text{ mol}$

The strong base (NaOH) reacts with the weak acid (NH₄⁺) to form a weak base (NH₃):

	NH ₄ Cl	NaOH	NH ₃
Before	0.2 mol	$2.0 \times 10^{-4} \text{ mol}$	0.2 mol
After	0.19998 mol	0	0.2002 mol

The small change in $[A^-]/[HA]$ means that the pH only goes up slightly. In the starting buffer, $[A^-]/[HA] = 1.00$, and $\text{pH} = \text{p}K_a$. After adding base, $[A^-]/[HA] = 1.002$, meaning that pH increased by $\log(1.002) = 0.001$ units.

5. Which of the following is/are the **conjugate acid of a strong base**?

(i) NH_4^+ (ii) H_2O (iii) H_3O^+ (iv) H_2 (v) CH_3COOH

- A) ii, iii
 B) i, iii, v
 C) i, v
D) ii, iv
 E) ii (1 mark)

	<u>conjugate base</u>	<u>weak or strong?</u>
(i)	NH_3	weak
(ii)	OH^-	strong
(iii)	H_2O	weak
(iv)	H^+	strong
(v)	CH_3COO^-	weak

6. **Which indicator** is best suited to determine the equivalence point of a titration of diethylamine with HBr ?

Data: $\text{p}K_b$ (diethylamine) = 3.16

- A) Bromocresol blue ($\text{p}K_{\text{HIN}} \approx 4$)**
 B) Methyl Violet ($\text{p}K_{\text{HIN}} \approx 1$)
 C) Phenolphthalein ($\text{p}K_{\text{HIN}} \approx 9$)
 D) Phenol Red ($\text{p}K_{\text{HIN}} \approx 7.5$)
 E) Alizarin Yellow-R ($\text{p}K_{\text{HIN}} \approx 11$)

In the titration of a weak base (diethylamine) with a strong acid (HBr), the pH goes from slightly basic, through the buffer region, to the equivalence point (mol strong acid added = mol weak base). At the equivalence point, the weak base has been converted into a weak acid, and the solution is that of a weak acid salt. Thus, pH at the equivalence point is < 7 , meaning that an indicator with $\text{p}K_{\text{HIN}}$ slightly below 7 would be best - in this case, bromocresol blue.

7. Which of the following statements about buffers is **FALSE**?

- A) Buffers resist changes in pH from the addition of limiting amounts of strong acid or strong base.
 B) The percent ionization of the weak acid in a buffer system is less than the percent ionization of the weak acid on its own because of the common ion effect.
C) The buffer capacity for a weak acid (HA) / weak base (A^-) system is at a maximum when $[\text{A}^-]$ and $[\text{HA}]$ are in ratios of 1:10 or 10:1.
 D) Buffer range is the range of pH values over which a buffer solution resists changes to pH, and the most effective range is $\text{pH} = \text{p}K_a \pm 1$.
 E) Adding a small amount of water to a buffer does not change the pH because the ratio of weak acid and weak base remains the same.

The buffer capacity is maximal when $\text{pH} = \text{p}K_a$, and therefore $[\text{A}^-] = [\text{HA}]$. The buffer capacity is decreased when $[\text{A}^-]/[\text{HA}] = 0.1$ or 10, as the buffer is at the very limit of its effective range. A small addition of base or acid will push it out of the buffer range.

8. A student obtains an unknown indicator which is colourless in neutral water. When placed in a 0.1 M solution of HCl or NH_4Cl , the indicator remains colourless. When placed in 0.1 M solution of NaNO_2 , the indicator turns blue. What is the **approximate** $\text{p}K_a$ of the indicator?
- A) 13.0 - 14.0
 B) 1.0 - 2.0
 C) 8.0 - 9.0
 D) 2.0 - 4.0
 E) 5.0 - 6.0

The indicator is colourless in neutral water (pH 7), 0.1 M HCl (pH 1.0) and 0.1 M NH_4Cl (pH < 7). It turns blue in the presence of a weak base (pH > 7). The colour change must come at pH > 7, so the answer is not B, D, or E. A is a strongly basic solution (equivalent to 0.1 - 1.0 M NaOH), so the answer is C.

9. In the titration of a weak acid with a strong base, the solution at the **equivalence point** is?
- A) neutral
 B) strongly basic
 C) weakly acidic
 D) strongly acidic
 E) weakly basic

In the titration of a weak acid with a strong base, the pH goes from slightly acidic, through the buffer region, to the equivalence point (mol strong base added = mol weak acid). At the equivalence point, the weak acid has been converted into a weak base, and the solution is that of a weak basic salt. Thus, pH at the equivalence point is > 7 and weakly basic.

10. What is the **pH** of the solution that results when the following substances are added to water to create 1.00 L of solution?
 Data: $K_b(\text{CH}_3\text{COO}^-) = 5.6 \times 10^{-10}$

0.100 mol HCl
 0.250 mol NaCH_3COO
 0.100 mol NaCl

- A) 5.13
 B) 1.00
 C) 5.15
 D) 4.92
 E) 3.26

0.100 mol HCl - strong acid
 0.250 mol NaCH_3COO - weak base
 0.100 mol NaCl - neutral salt, ignore it
 $K_b(\text{CH}_3\text{COO}^-) = 5.6 \times 10^{-10}$, therefore $\text{p}K_b = 9.25$
 $\text{p}K_a(\text{CH}_3\text{COOH}) = \text{p}K_w - \text{p}K_b = 14 - 9.25 = 4.75$

	HCl	NaCH ₃ COO	CH ₃ COOH
Before	0.100 mol	0.250 mol	0
After	0	0.150 mol	0.100 mol

Henderson-Hasselbalch equation:

$$\begin{aligned}\text{pH} &= \text{p}K_a + \log([\text{CH}_3\text{COO}^-]/[\text{CH}_3\text{COOH}]) \\ &= 4.75 + \log(0.150/0.100) \\ &= 4.92\end{aligned}$$

11. Indicate the **pH** of a 0.020 M solution of benzoic acid in water.

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

- A) 3.12
B) 2.95
 C) 2.85
 D) 4.50
 E) 2.61

A weak acid solution requires using an ICE table and small x approximation:

HA = benzoic acid

A⁻ = benzoate

	HA +	H ₂ O	\rightleftharpoons	A ⁻	+ H ₃ O ⁺
I	0.020 M	-		0	0
C	-x M	-		+x M	+x M
E	0.020 - x M	-		x M	x M

$$K_a = [\text{A}^-][\text{H}_3\text{O}^+]/[\text{HA}] = x^2/[\text{HA}], \text{ therefore } x = \sqrt{\{K_a \times [\text{HA}]\}}$$

$$\begin{aligned}x &= \sqrt{\{(6.3 \times 10^{-5}) \times (0.02)\}} \\ &= 1.12 \times 10^{-3} \text{ M}\end{aligned}$$

therefore:

$$[\text{H}_3\text{O}^+] = 1.12 \times 10^{-3} \text{ and pH} = 2.95.$$

12. Acetic acid (10.00 mL) is added to a beaker and diluted with ≈ 40 mL of distilled water before starting the titration with NaOH. The addition of ≈ 40 mL of distilled water **will change** which of the following:

- A) The pH of the NaOH solution
 B) The volume of NaOH required to reach the maximum buffer capacity
C) The initial pH of the acetic acid solution
 D) The volume of NaOH needed to reach the equivalence point
 E) The pH at the 1/2 equivalence point

The pH of a weak acid solution depends on the acid's concentration, therefore diluting the solution changes (increases) the pH.

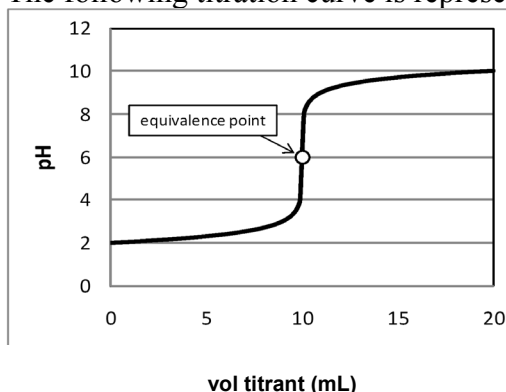
13. Which of the following acids would be the best choice for the weak acid component of a buffer with pH 6.9?

- A) $\text{C}_6\text{H}_5\text{COOH}$ (benzoic acid) $K_a = 6.3 \times 10^{-5}$
 B) HCN $K_a = 6.2 \times 10^{-10}$
 C) H_3PO_4 $K_a = 7.1 \times 10^{-3}$
 D) HOCl $K_a = 2.9 \times 10^{-8}$
 E) $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$ (citric acid) $K_a = 7.5 \times 10^{-4}$

HOCl is the only weak acid with a $\text{p}K_a = \text{pH} \pm 1$ of the desired $\text{pH} = 6.9$.

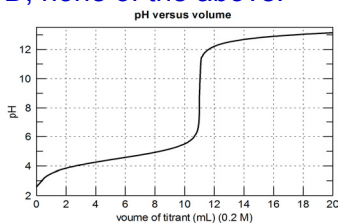
A) $\text{C}_6\text{H}_5\text{COOH}$ (benzoic acid)	$K_a = 6.3 \times 10^{-5}$	$\text{p}K_a = 4.20$
B) HCN	$K_a = 6.2 \times 10^{-10}$	$\text{p}K_a = 9.21$
C) H_3PO_4	$K_a = 7.1 \times 10^{-3}$	$\text{p}K_a = 2.15$
D) HOCl	$K_a = 2.9 \times 10^{-8}$	$\text{p}K_a = 7.54$
E) $\text{H}_3\text{C}_6\text{H}_5\text{O}_7$ (citric acid)	$K_a = 7.5 \times 10^{-4}$	$\text{p}K_a = 3.12$

14. The following titration curve is representative of what type of titration?



- A) a strong base (beaker) with a weak acid (buret) titration
 B) a strong acid (beaker) with strong base (buret) titration
 C) a weak acid (beaker) with strong base (buret) titration
 D) the graph does not represent any of the above titrations
 E) a weak base (beaker) with strong acid (buret) titration

pH increases with increasing volume of titrant, therefore the titrant is a strong base (not A or E). The equivalence point is at pH 6, so it is not a strong acid/strong base titration (not B), as it would have $\text{pH} = 7$ at the equivalence point. There is no buffer region (compare to the titration below), therefore it is not a weak acid/strong base titration (not C). The only remaining option is D, none of the above.



15. Given the reaction $A \rightarrow 2G$, and the following time versus $[G]$ data:

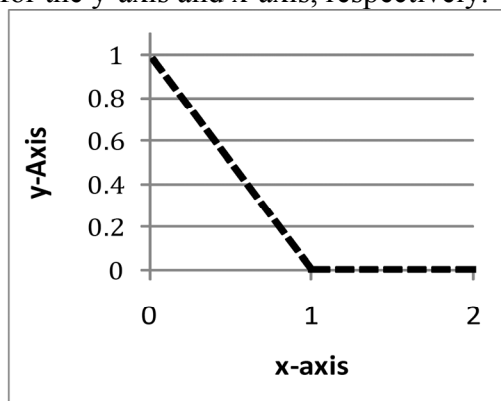
time (s)	$[G]$ (M)
0	0
18	0.032
24	0.037

Which of the following statements is **TRUE**?

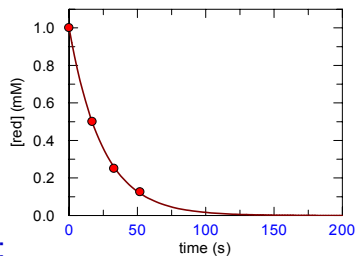
- A) The average reaction rate over 18 s is 1.8×10^{-3} M/s.
 B) The average reaction rate over 24 s is 7.7×10^{-4} M/s.
 C) $v_0 = 1.8 \times 10^{-3}$ M/s
 D) The reaction is zero order with respect to A.
 E) $-\Delta[A]/\Delta t = \Delta[G]/\Delta t$

- A) average rate = $\frac{1}{2}\Delta[G]/\Delta t = 0.5 \times 0.032 \text{ M} / 18 \text{ s} = 8.9 \times 10^{-4} \text{ M/s}$.
 B) average rate = $\frac{1}{2}\Delta[G]/\Delta t = 0.5 \times 0.037 \text{ M} / 24 \text{ s} = 7.7 \times 10^{-4} \text{ M/s}$.
 C) This is not a zero order reaction because the average rate over 18 s (see A) is faster than over 24 s (see B). Without knowing the rate law, it is not possible to calculate v_0 .
 D) The reaction rate changes with time (see C), therefore it is not zero order.
 E) It is necessary to take the stoichiometry into account: $-\Delta[A]/\Delta t = \frac{1}{2}\Delta[G]/\Delta t$.

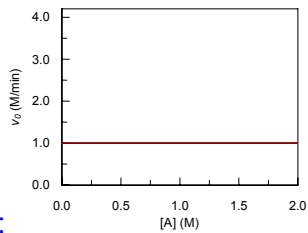
16. The graph below describes the conversion of A into G. Choose the **correct axis labels** for the y-axis and x-axis, respectively:



- A) $[A]$, and time (s), for a first order reaction
 B) rate (M s^{-1}), and $[G]$, for a reversible reaction
 C) rate (M s^{-1}), and time (s), for a zero order reaction
 D) $[A]$, and time (s), for a zero order reaction
 E) k ($\text{M}^{-1}\text{s}^{-1}$), and time (s), for a first order reaction



- A) The graph would look like this:
- B) Depends on the reaction order w.r.t. $[G]$, but it could not be a straight line with negative slope.



- C) The graph would look like this:
- D) Correct.
- E) k is constant for any kind of reaction, therefore it would look like (C).

17. The reaction $2 \text{H}_2\text{O}_2 \rightarrow 2 \text{H}_2\text{O} + \text{O}_2$ produces 3.2 g/min of O_2 from a 1 L reaction volume. Which of the following statements is **FALSE**?

- A) $-\Delta[\text{H}_2\text{O}_2]/\Delta t = \Delta[\text{H}_2\text{O}]/\Delta t$
- B) $-\Delta[\text{H}_2\text{O}_2]/\Delta t = 0.2 \text{ M/min}$
- C) $v_0 = 0.2 \text{ M/min}$
- D) $\Delta[\text{O}_2]/\Delta t = 0.1 \text{ M/min}$
- E) $-\frac{1}{2}d[\text{H}_2\text{O}_2]/dt = d[\text{O}_2]/dt$

The reaction produces 0.1 M/min ($3.2 \text{ g/min} \div 32 \text{ g/mol} \div 1 \text{ L}$) of O_2 .

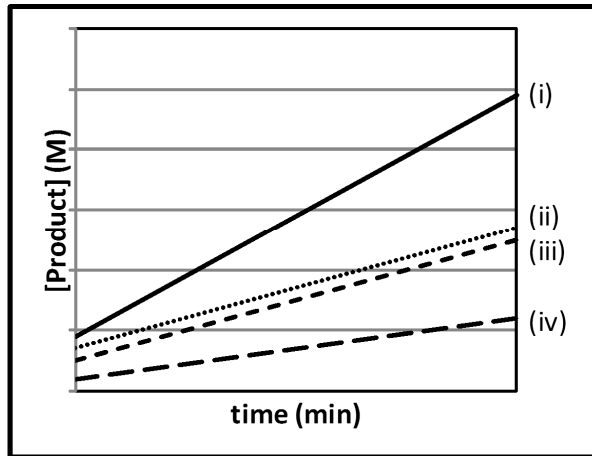
18. Based on the observed rates, what is the **rate law** for the reaction $A + B \rightarrow G$?

[A] (M)	[B] (M)	v_0 (M/s)
0.1	0.1	0.4
0.1	0.4	1.6
0.2	0.2	0.8

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

Compared with the first set of conditions, increasing [B] 4-fold in the second line increases v_0 4-fold. Therefore the reaction is first order w.r.t. B. In the last line, doubling [B] should double v_0 , which it does. Therefore the fact that doubling [A] in the last line had no effect on v_0 means that the reaction was zero order w.r.t. A. Therefore: $v_0 = k[A]^0[B]$ (or $v_0 = k[B]$).

19. The following [product] versus time data were obtained for the reaction $A + B + C \rightarrow G$. What is the **rate law** for this reaction?



	[A] (M)	[B] (M)	[C] (M)
(i)	0.2	0.2	0.2
(ii)	0.2	0.1	0.1
(iii)	0.2	0.2	0.1
(iv)	0.1	0.1	0.1

- A) $v_0 = k[A][B][C]$
 B) $v_0 = k[A]^2[B]^{-1}[C]$
C) $v_0 = k[A][C]$
 D) $v_0 = k[A][B][C]^2$
 E) $v_0 = k[A][C]^2$

Starting from (iv), we see in (ii) that doubling [A] doubles the rate, thus the reaction is first order w.r.t. A. In (iii), [B] is doubled compared to (ii), but the rate does not change, therefore the reaction is zero order w.r.t. B. The rate in (i) is double (iii) and [C] is doubled, therefore the reaction is first order w.r.t. C. Therefore: $v_0 = k[A][C]$.

20. A first order reaction has $k = 0.0112 \text{ s}^{-1}$. What is the **half-life (in s)**?

- A) 81.9
B) 61.9
 C) 78.3
 D) 53.2
 E) 48.6

$$t_{1/2} = \ln(2)/k$$

therefore:

$$t_{1/2} = \ln(2)/(0.0112 \text{ s}^{-1}) = 61.9 \text{ s}$$

Section #2 – These questions are worth three marks each.

21. In the first order decomposition reaction of compound A, [A] was 0.1 M at $t = 450$ s after the reaction started, and 0.050 M at $t = 750$ s after the reaction started. Indicate the correct $[A]_0$:

- A) 0.31 M
- B) 0.22 M
- C) 0.20 M
- D) 0.15 M
- E) 0.28 M

There are several ways to solve this question. The simplest is to recognize that [A] decreased by half in the interval from 450 s to 750 s, therefore $t_{1/2} = 300$ s.

We know that

$$k = \ln(2)/t_{1/2},$$

therefore,

$$k = 2.31 \times 10^{-3} \text{ s}^{-1}$$

Rearranging the integrated rate equation gives:

$$[A]_0 = [A]_t / e^{-kt}$$

Plugging in the values from either time point should give the same value of $[A]_0$.

$$\begin{aligned} [A]_0 &= 0.1 \text{ M} / \exp(-2.31 \times 10^{-3} \text{ s}^{-1} \times 450 \text{ s}) \\ &= 0.28 \text{ M} \end{aligned}$$

22. In the titration of 10.00 mL of 0.2134 M NaOH with 0.1345 M HCl, at **what volume of HCl** added would the pH be 11.668?

- A) 13.5
- B) 15.0**
- C) 7.44
- D) 19.1
- E) 17.8

The key solving this question is recognizing that $[\text{NaOH}]$ is decreasing due both to HCl addition, and due to dilution as the titration proceeds. We will set up an equation to calculate $[\text{OH}^-]$ as a function of the volume of HCl added, then plug in our target $[\text{OH}^-]$ needed to give pH 11.668.

Our target $[\text{NaOH}]$ is given by:

$$\text{pOH} = 14 - \text{pH} = 2.332$$

and

$$[\text{OH}^-] = 10^{-2.332} = 4.656 \times 10^{-3} \text{ M} \quad (1)$$

$$n(\text{OH}^-)_{\text{initial}} = 0.2134 \text{ M} \times 0.01000 \text{ L} = 2.134 \times 10^{-3} \text{ mol} \quad (2)$$

$$\text{vol}(\text{NaOH})_{\text{initial}} = 0.01000 \text{ L} \quad (3)$$

The number of moles of OH^- during the titration is:

$$\begin{aligned} n(\text{OH}^-) &= n(\text{OH}^-)_{\text{initial}} - n(\text{HCl})_{\text{added}} \\ &= n(\text{OH}^-)_{\text{initial}} - 0.1345 \text{ M} \times \text{vol}(\text{HCl}) \end{aligned}$$

The volume is:

$$\text{vol} = \text{vol}(\text{NaOH})_{\text{initial}} + \text{vol}(\text{HCl})$$

Therefore,

$$[\text{OH}^-] = \{n(\text{OH}^-)_{\text{initial}} - 0.1345 \times \text{vol}(\text{HCl})\} / \{\text{vol}(\text{NaOH})_{\text{initial}} + \text{vol}(\text{HCl})\}$$

$$[\text{OH}^-] \times \{\text{vol}(\text{NaOH})_{\text{initial}} + \text{vol}(\text{HCl})\} = \{n(\text{OH}^-)_{\text{initial}} - 0.1345 \times \text{vol}(\text{HCl})\}$$

$$\text{vol}(\text{HCl}) \times \{[\text{OH}^-] + 0.1345\} = n(\text{OH}^-)_{\text{initial}} - \text{vol}(\text{NaOH})_{\text{initial}} \times [\text{OH}^-]$$

$$\text{vol}(\text{HCl}) = \{n(\text{OH}^-)_{\text{initial}} - \text{vol}(\text{NaOH})_{\text{initial}} \times [\text{OH}^-]\} / \{[\text{OH}^-] + 0.1345\}$$

Given our target $[\text{OH}^-]$ (1) and values (2) and (3):

$$\text{vol}(\text{HCl}) = (2.134 \times 10^{-3} \text{ mol} - 0.0100 \text{ L} \times 4.656 \times 10^{-3} \text{ M}) / (4.656 \times 10^{-3} \text{ M} + 0.1345 \text{ M})$$

$$\text{vol}(\text{HCl}) = 0.0150 \text{ L}$$

(The quick, approximate method is to recognize that at pH 11.668, the NaOH has been almost completely neutralized by HCl, so the volume of HCl added will be a little less than the equivalence point. The volume of HCl at the equivalence point would be $10.00 \text{ mL} \times (0.2134 \text{ M} / 0.1345 \text{ M}) = 15.9 \text{ mL}$, which is a little more than 15.0 mL.)

23. What is the **pH** of a solution resulting from the titration to equivalence of 10.0 mL of HF ($pK_a = 3.18$) with 22.3 mL of 0.1289 M KOH?

- A) 11.48
B) 8.06
 C) 9.51
 D) 7.22
 E) 3.21

At the equivalence point, HF is converted completely to KF, and its pH will depend on its concentration. We don't know $[HF]_{\text{initial}}$, but we do know the number of moles of KOH at the equivalence point.

$$\begin{aligned} n(\text{KOH}) &= \text{vol}(\text{KOH}) \times [\text{KOH}] \\ &= (0.0223 \text{ L} \times 0.1289 \text{ M}) \\ &= 2.87 \times 10^{-3} \text{ mol} \end{aligned}$$

At the equivalence point,

$$\begin{aligned} [\text{KF}] &= n(\text{KOH}) / (\text{vol}(\text{KOH}) + \text{vol}(\text{HF})) \\ &= 2.87 \times 10^{-3} \text{ mol} / (0.0223 \text{ L} + 0.0100 \text{ L}) \\ &= 0.0890 \text{ M} \end{aligned}$$

If $pK_a = 3.18$ for HF, then:

$$\begin{aligned} pK_b &= 14 - 3.18 \\ &= 10.82 \end{aligned}$$

and

$$K_b = 10^{-10.82} = 1.51 \times 10^{-11}$$

Use an ICE table to solve for $[\text{OH}^-]$ of a weak base:

$$\begin{aligned} \text{A}^- &= \text{KF} \\ \text{HA} &= \text{HF} \end{aligned}$$

	A^-	+	H_2O	\rightleftharpoons	HA	+	OH^-
I	0.0890 M		-		0		0
C	-x M				+x M		
E	0.0890 - x M		-		x M		x M

$$K_b = [\text{HA}][\text{OH}^-]/[\text{A}^-] = x^2/[\text{A}^-], \text{ therefore } x = \sqrt{\{K_b \times [\text{A}^-]\}}$$

$$\begin{aligned} x &= \sqrt{\{(1.51 \times 10^{-11}) \times (0.0890)\}} \\ &= 1.16 \times 10^{-6} \text{ M} \end{aligned}$$

therefore:

$$[\text{OH}^-] = 1.16 \times 10^{-6}, \text{ pOH} = 5.94$$

therefore:

$$\begin{aligned} \text{pH} &= 14 - 5.94 \\ &= 8.06 \end{aligned}$$

24. Consider the table below, showing the decrease of [A] during the reaction $A \rightarrow G$. Indicate the **FALSE** statement(s):

time (h)	[A] (M)
1.0	1.60
2.0	1.20
3.0	0.80
4.0	0.40

- (i) $k = 1.11 \times 10^{-4} \text{ M s}^{-1}$
 (ii) $k = 4.00 \times 10^{-1} \text{ s}^{-1}$
 (iii) $[A]_0 = 2.40 \text{ M}$

- A) i, iii
 B) iii (1 mark)
 C) ii, iii
 D) i, ii
 E) i

The rate does not change during the course of the reaction ($\Delta[A]/\Delta t = -0.4 \text{ M/h}$), therefore it is zero order w.r.t. A.

$$v = 0.4 \text{ M/h,}$$

or

$$v = 1.11 \times 10^{-4} \text{ M/s, therefore:}$$

- (i) is true
 (ii) is false because the number is wrong (see (i)) and the units are incorrect for a zero order reaction
 (iii) is false, $[A]_0 = 2.00 \text{ M}$.

25. You are given 250.0 mL of 0.125 M propanoic acid ($\text{CH}_3\text{CH}_2\text{COOH}$), and you want to adjust its pH to 4.25 by adding 1.00 M $\text{NaCH}_3\text{CH}_2\text{COO}(\text{aq})$. **What volume (mL)** of the $\text{NaCH}_3\text{CH}_2\text{COO}$ solution should be added?

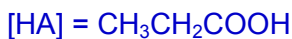
Data: $K_a(\text{CH}_3\text{CH}_2\text{COOH}) = 1.35 \times 10^{-5}$

- A) 9.56 mL
- B) 16.2 mL
- C) 11.4 mL
- D) 7.50 mL**
- E) 3.25 mL

Target pH = 4.25.

$$\begin{aligned} \text{p}K_a &= -\log(1.35 \times 10^{-5}) \\ &= 4.87 \end{aligned}$$

From the Henderson-Hasselbalch equation:



$$\begin{aligned} [\text{A}^-]/[\text{HA}] &= 10^{(\text{pH} - \text{p}K_a)} \\ &= 0.240 \end{aligned}$$

$$[\text{A}^-] = 0.240 \times [\text{HA}]$$

From the initial conditions, the number of moles of HA is:

$$\begin{aligned} n(\text{HA}) &= (0.2500 \text{ L} \times 0.125 \text{ M}) \\ &= 0.03125 \text{ mol} \end{aligned}$$

Because we are making a buffer, concentrations are not important, only the ratio of HA and A^- :

$$\begin{aligned} n(\text{A}^-) &= 0.240 \times n(\text{HA}) \\ &= 0.240 \times 0.03125 \text{ mol} \\ &= 7.50 \times 10^{-3} \text{ mol} \end{aligned}$$

$$\begin{aligned} \text{vol}(\text{A}^-) &= n(\text{A}^-) / [\text{A}^-] \\ &= 7.50 \times 10^{-3} \text{ mol} / 1 \text{ M} \\ &= 7.50 \times 10^{-3} \text{ L or } 7.50 \text{ mL} \end{aligned}$$

Answer key: Version 1

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

Answer key: Version 2

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

Answer key: Version 3

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

Answer key: Version 4

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