

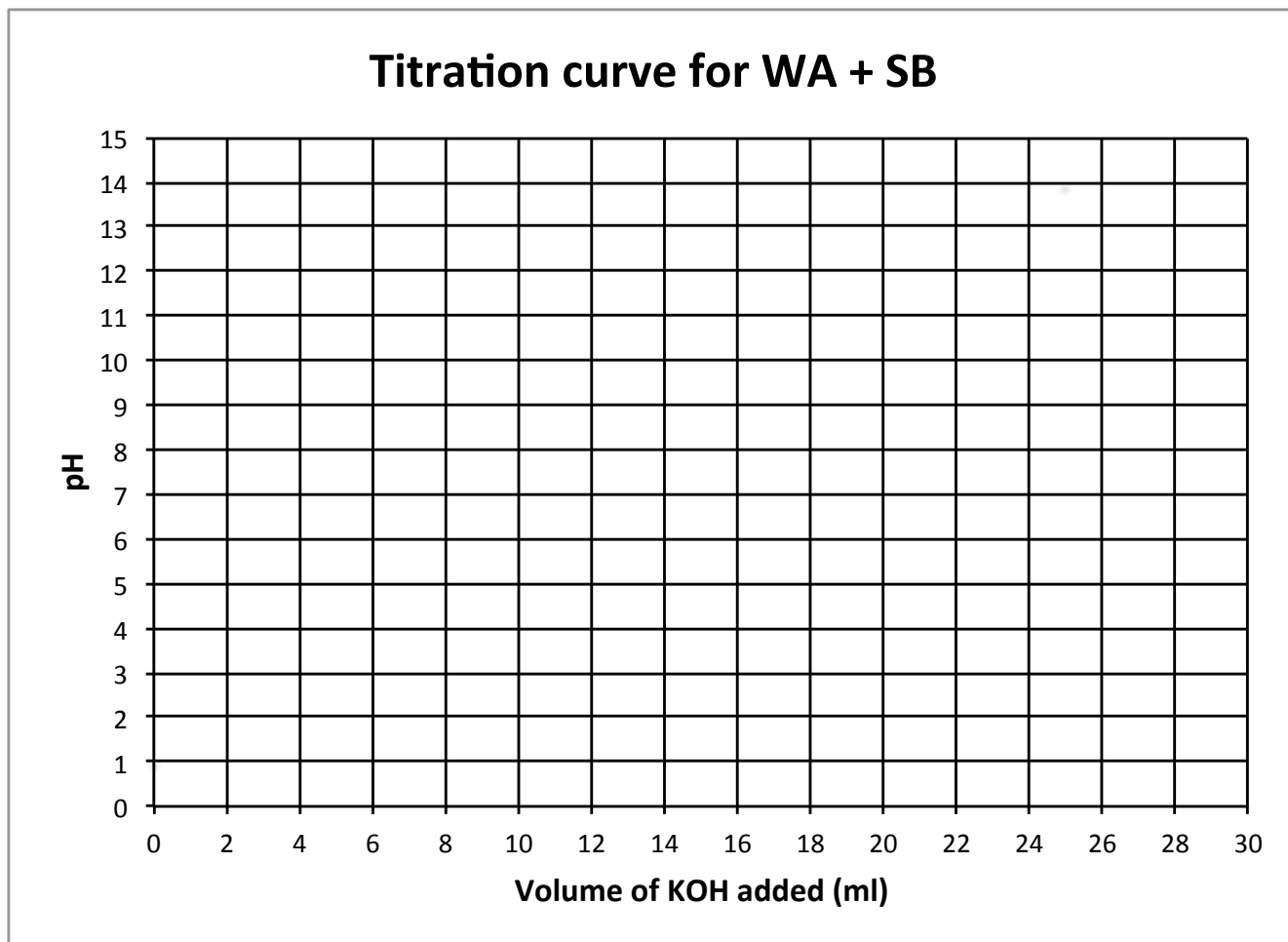
P G4 A (pg 1 of 4) **Acid – Base Titration Calculations (WA + SB)**

Name _____ Per _____

Consider the changes (pH, $[H^+]$, $[OH^-]$) when titrating 50.0 ml of 0.20 M solution of lactic acid, $HC_3H_5O_3$ with 0.40 M potassium hydroxide. Then construct a titration curve on the graph on page 2. The K_a of lactic acid = 1.4×10^{-4}

1. Calculate the pH of 50.0 ml of a 0.20 M solution of lactic acid, $HC_3H_5O_3$.
2. What volume of 0.40 M KOH would be needed to neutralize this 50.0 ml of a 0.20 M solution of lactic acid, $HC_3H_5O_3$? (This should be the first step you would take to solve any titration problem, even if you weren't asked to do it.)
3. Calculate the pH of 50.0 ml of a 0.20 M solution of lactic acid, $HC_3H_5O_3$ after it has been titrated with 10.0 ml of 0.40 M KOH?
4. Calculate the pH of 50.0 ml of a 0.20 M solution of lactic acid, $HC_3H_5O_3$ after it has been titrated with a total of 12.5 ml of 0.40 M KOH?
5. Calculate the pH of 50.0 ml of a 0.20 M solution of lactic acid, $HC_3H_5O_3$ after it has been titrated with a total of 20.0 ml of 0.40 M KOH?
6. Calculate the pH of 50.0 ml of a 0.20 M solution of lactic acid, $HC_3H_5O_3$ after it has been titrated with a total of 25.0 ml of 0.40 M KOH?
7. Calculate the pH of 50.0 ml of a 0.20 M solution of lactic acid, $HC_3H_5O_3$ after it has been titrated with a total of 30.0 ml of 0.40 M KOH?

Make a sketch of the titration curve using the volumes and pH values calculated for the titration of 50 ml of 0.20 M lactic acid with 0.40 M KOH. Placing a point on the graph for problems 3 – 7 from the previous page and #'s 8 & 9 below.



As you construct this graph, you will get a much more accurate sense of the curve if you make two more calculations and place two more points on the graph.

8. Calculate the pH of 50.0 ml of a 0.20 M solution of lactic acid, $\text{HC}_3\text{H}_5\text{O}_3$ after it has been titrated with a total of 24.0 ml of 0.40 M KOH?

9. Calculate the pH of 50.0 ml of a 0.20 M solution of lactic acid, $\text{HC}_3\text{H}_5\text{O}_3$ after it has been titrated with a total of 26.0 ml of 0.40 M KOH?

Be sure that if you need help, you refer to NS G.1 – Acid-Base Titrations.

First you need to recognize that this is a WA being titrated with a SB.

- $K_a = \frac{[H^+][A^-]}{[HA]}$ This is an “x²” problem. $1.4 \times 10^{-4} = \frac{x^2}{[0.20]}$ $x = [H^+] = 5.3 \times 10^{-3} \text{ M}$ so **pH = 2.28**
- It is ALWAYS important to calculate the volume required to reach the equivalence point; moles of acid = moles of base, $M_a V_a = M_b V_b$
 - so $0.20 \text{ M} \times 50 \text{ ml} = 0.40 \text{ M} \times V_b$
 - thus the volume of strong base needed to reach the equivalence point is **25.0 ml of KOH**
- At this point you are only partway through the titration, but not yet to the halfway point.
 - Calculate the millimoles of lactic acid ($0.2\text{M} \times 50\text{ml}$) and millimoles of incoming strong base ($0.4\text{M} \times 10\text{ml}$)
 - Subtract the incoming strong base from the weak acid, to find out the remaining un-neutralized weak acid (6 mmole) and there so very little A⁻ in solution to start but the incoming strong base will drive the equilibrium to the right and produce an amount of mmoles of A⁻ equivalent to the incoming strong base (4 mmole) so you will add that to the conjugate weak base (which was originally ~ 0)

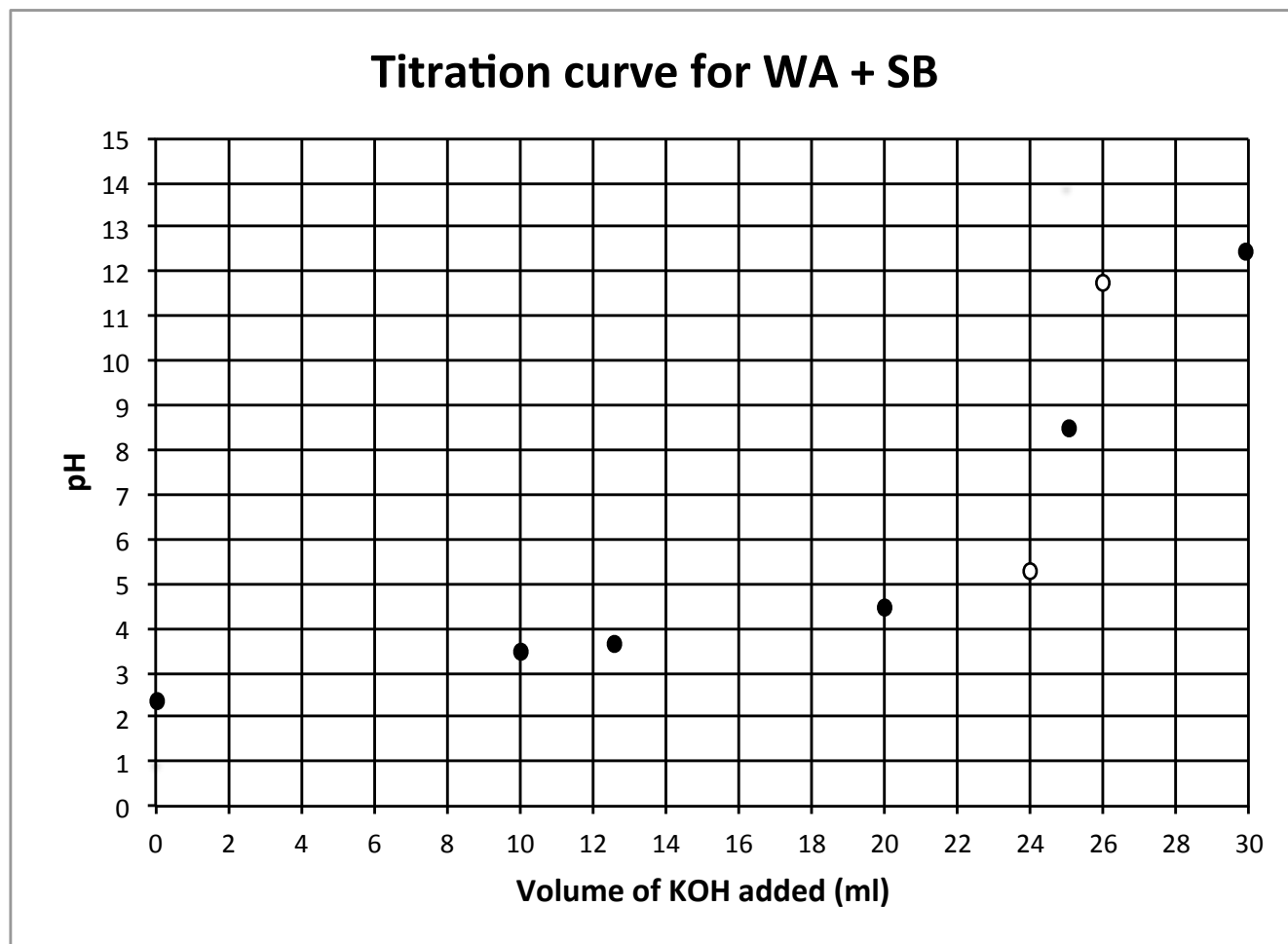
$$K_a = \frac{[H^+][A^- \text{ -incomingSA} + \text{incomingSB}]}{[HA \text{ +incomingSA} - \text{incomingSB}]} \quad K_a = \frac{[H^+](A^- + \text{moleOH}^-)}{(HA - \text{moleOH}^-)} \quad K_a = \frac{[H^+](\sim 0 + 4\text{moleOH}^-)}{(\sim 10\text{mmol} - 4\text{moleOH}^-)} \quad 1.4 \times 10^{-4} = \frac{[H^+]4\text{mmol}}{6\text{mmol}}$$

- Solve and $[H^+] = 2.1 \times 10^{-4} \text{ M}$ so **pH = 3.68**
- Yippee! This is the halfway point of the titration, because half of the weak acid is reacted (10 mmol – 5 mmol), and half of the conjugate weak base forms (5 mmol) $K_a = \frac{[H^+](\sim 0 + 5\text{moleOH}^-)}{(\sim 10\text{mmol} - 5\text{moleOH}^-)}$ and thus $[HA] = [A^-]$ resulting in $K_a = [H^+] \left(\frac{A^-}{HA} = \right)$
 - so always at the halfway point, the $K_a = [H^+] = 1.4 \times 10^{-4}$ and **pH = 3.85**
 - At this point you are again only partway through the titration.
 - Calculate the millimoles of lactic acid ($0.2 \text{ M} \times 50 \text{ ml}$) and millimoles of incoming strong base ($0.4\text{M} \times 20\text{ml}$)
 - Subtract the incoming strong base from the weak acid, to find out the remaining un-neutralized acid (2 mmole) and there is so very little A⁻ in solution to start, ~0, but the incoming base will drive the equilibrium to the right and produce an amount of mmoles of A⁻ equivalent to the incoming strong base (8 mmole) so you will add that to the conjugate weak base (which was originally ~ 0)



$$K_a = \frac{[H^+][A^- \text{ -incomingSA} + \text{incomingSB}]}{[HA \text{ +incomingSA} - \text{incomingSB}]} \quad K_a = \frac{[H^+](A^- + \text{moleOH}^-)}{(HA - \text{moleOH}^-)} \quad K_a = \frac{[H^+](\sim 0 + 8\text{moleOH}^-)}{(\sim 10\text{mmol} - 8\text{moleOH}^-)} \quad 1.4 \times 10^{-4} = \frac{[H^+]8\text{mmol}}{2\text{mmol}}$$

- Solve $[H^+] = 3.5 \times 10^{-5} \text{ M}$ so **pH = 4.46**
- At this point you are at the equivalence point and all of the WA has been neutralized – the equilibrium has been driven all the way to the right, but the pH is NOT 7 because there will be conjugate WB, (A⁻) in the solution which does not just “sit there,” but hydrolyzes with water causing the pH to be above 7. *So here you must switch gears.*
 - You have already calculated the millimoles of WA, and that will be the amount of conjugate WB, A⁻ in solution because the incoming strong base has driven the neutralization “all the way” and all the WA has been converted to WB.
 - thus $[A^-] = 10 \text{ mmole}$ divided by total volume (75 ml) = 0.133 M
 - You can then use the $[A^-]$ in a simple “x²” WB calculation to determine the $x = [OH^-]$
 - $K_b = \frac{[OH^-][HA]}{[A^-]}$ (Of course you need to calculate the K_b from the K_a remember: $K_a \times K_b = 1 \times 10^{-14}$)
 - thus $K_b = 7.1 \times 10^{-11} = \frac{x^2}{[0.133]}$ and $x = [OH^-] = 3.1 \times 10^{-6} \text{ M}$
 - After solving for $x = [OH^-]$ you can $-\log[OH^-]$ determine the pOH = 5.51 the “14-it” for **pH = 8.49**
 - At this point in the titration the amount of strong base added exceeds the amount of weak acid.
 - You know that the equivalence point occurs at 25 ml of KOH added, thus 30 ml of KOH is 5 ml beyond the equivalence point resulting in an extra 2 mmole of strong base (since $0.40 \text{ M} \times 5 \text{ ml}$ extra base = 2 mmol extra SB) (OR 10mm WA with 12 mmol SB = 2 mm extra SB)
 - Divide this extra 2 mmole of strong base by total volume (80 ml) to get 0.025 M of strong base in excess. (The rest of the incoming strong base was neutralized by H⁺ put out by the dissociating WA.) This SB which contributes far more OH⁻ than the conjugate weak base does, so you need not concern yourself with any minute “extra” OH⁻ contributed by the WB.
 - Thus $-\log[0.025] = \text{pOH} = 1.60$, thus “14-it” to get **pH = 12.40**



8. At this point, just like the calculation that you made in #5, you are again only partway through the titration.
- Calculate the millimoles of lactic acid ($0.2 \text{ M} \times 50 \text{ ml}$) and millimoles of incoming strong base ($0.4 \text{ M} \times 24 \text{ ml}$)
 - Subtract the incoming strong base from the weak acid, to find out the remaining un-neutralized weak acid (0.4 mmole). Realize that there so very little A^- in solution to start but the incoming strong base will drive the equilibrium to the right and produce an amount of mmoles of A^- equivalent to the incoming SB (9.6 mmole) so you will add that to the conjugate WB (which was originally ~ 0)

$$K_a = \frac{[H^+][A^- \text{ -incomingSA} + \text{incomingSB}]}{[HA \text{ +incomingSA} - \text{incomingSB}]} \quad K_a = \frac{[H^+](A^- + \text{moleOH}^-)}{(HA - \text{moleOH}^-)} \quad K_a = \frac{[H^+](\sim 0 + 9.6 \text{moleOH}^-)}{(\sim 10 \text{mmol} - 9.6 \text{moleOH}^-)} \quad 1.4 \times 10^{-4} = \frac{[H^+]9.6 \text{mmol}}{0.4 \text{mmol}}$$

- Solve $[H^+] = 5.8 \times 10^{-6} \text{ M}$ so **pH = 5.23**
9. At this point in the titration the amount of SB added exceeds the amount of WA, just like in #7.
- You know that the equivalence point occurs at 25 ml of KOH added, thus 26 ml of KOH is 1 ml beyond the equivalence point. From previous calculations, we know that the mmole of WA is ($0.2 \text{ M} \times 50 \text{ ml}$) 10 mmoles, and the mmol of incoming SB is more, 10.4 mmoles ($0.40 \text{ M} \times 26 \text{ ml SB}$)
 - Subtract to get 0.4 mmole of excess SB, then divide by total volume (76 ml) to get 0.0055 M of strong base in excess. (The rest of the incoming SB was neutralized by H^+ put out by the dissociating WA.) This SB which contributes far more OH^- than the conjugate weak base contributes, so you need not concern yourself with any minute “extra” OH^- contributed by the WB.
 - Thus $-\log(0.00525) = \text{pOH} = 2.28$, thus “14-it” to get **pH = 11.72**