

EQUILIBRIUM PROBLEMS
APlus Alabama Fall 2022 WORKSHOP

1. A flask containing 1.00 atm of NO₂ and 1.50 atm of N₂O₄ is allowed to proceed to equilibrium at a given temperature.



When equilibrium is established the partial pressure of NO₂ is 0.500 atm.

- a) Calculate the magnitude of the equilibrium constant for this reaction at this temperature.

$2\text{NO}_2(\text{g}) \rightleftharpoons \text{N}_2\text{O}_4(\text{g})$ <p>Initial 1.00 atm 1.50 atm</p> <p>Change</p> <p>Equilibrium 0.500 atm</p>	<p>Units are atm so calc K_p</p> $K_p = \frac{P(\text{N}_2\text{O}_4)}{P(\text{NO}_2)^2} = \frac{P(1.750)}{P(0.500)^2} = 7.0$
$2\text{NO}_2(\text{g}) \rightleftharpoons \text{N}_2\text{O}_4(\text{g})$ <p>Initial 1.00 atm 1.50 atm</p> <p>Change -0.500 +0.250</p> <p>Equilibrium 0.500 atm 1.75 atm</p>	

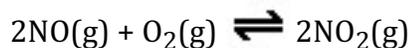
- b) After the reaction mixture has established equilibrium, 0.250 atm of N₂O₄ are added to the reaction flask. Which direction (reactant to products; product to reactants; or no change) will the reaction proceed to re-establish equilibrium? Explain in terms of Q and K_p.

$2\text{NO}_2(\text{g}) \rightleftharpoons \text{N}_2\text{O}_4(\text{g})$ <p>Equilibrium 0.500 atm 1.75 atm</p> <p>Add 0 0.250 atm</p> <p>New Initial 0.500 atm 2.00 atm</p>	<p>Units are atm so calc K_p</p> <p>K_p = 7.0</p> $Q = \frac{P(\text{N}_2\text{O}_4)}{P(\text{NO}_2)^2} = \frac{P(2.00)}{P(0.500)^2} = 8.0$
<p>Q > K_p To re-establish equilibrium the quotient must decrease. The reaction must proceed from products toward reactants</p>	

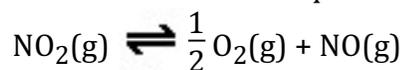
- c) Does the position of equilibrium favor the products side or the reactants side? Explain.

K_p is equal to 7.0, products are at higher pressure compared to reactants and the position of equilibrium favors the products.

2. The equilibrium constant for the reaction



At a certain temperature is 1.48×10^4 . The equilibrium constant for the reaction



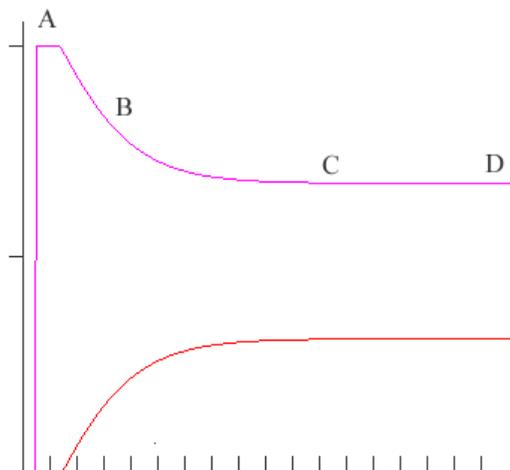
What is the equilibrium constant for this reaction at the same temperature?

We will assume the equilibrium constant is K_c

$$K_c = \frac{[\text{NO}_2]^2}{[\text{NO}]^2[\text{O}_2]} = 1.48 \times 10^4$$

$$\text{The new reaction has a } K = \frac{[\text{NO}][\text{O}_2]^{1/2}}{[\text{NO}_2]} = \frac{1}{(K_c)^{1/2}} = \frac{1}{(1.48 \times 10^4)^{1/2}} = 8.22 \times 10^{-3}$$

3. In the graph below the endothermic reaction $\text{BR}(g) \rightleftharpoons \text{B}(g) + \text{R}(g)$ is represented. Initially only $\text{BR}(g)$ is present in the reaction vessel. The marks along the x -axis are in 1-minute increments. The initial $[\text{BR}]$ (y -axis) is 2.0 M. The reaction begins about 1.5 minutes in this case.



- a) At what point (indicate a letter) does the reaction attain equilibrium? NOTE: You can also label the graph if you prefer.

At point C

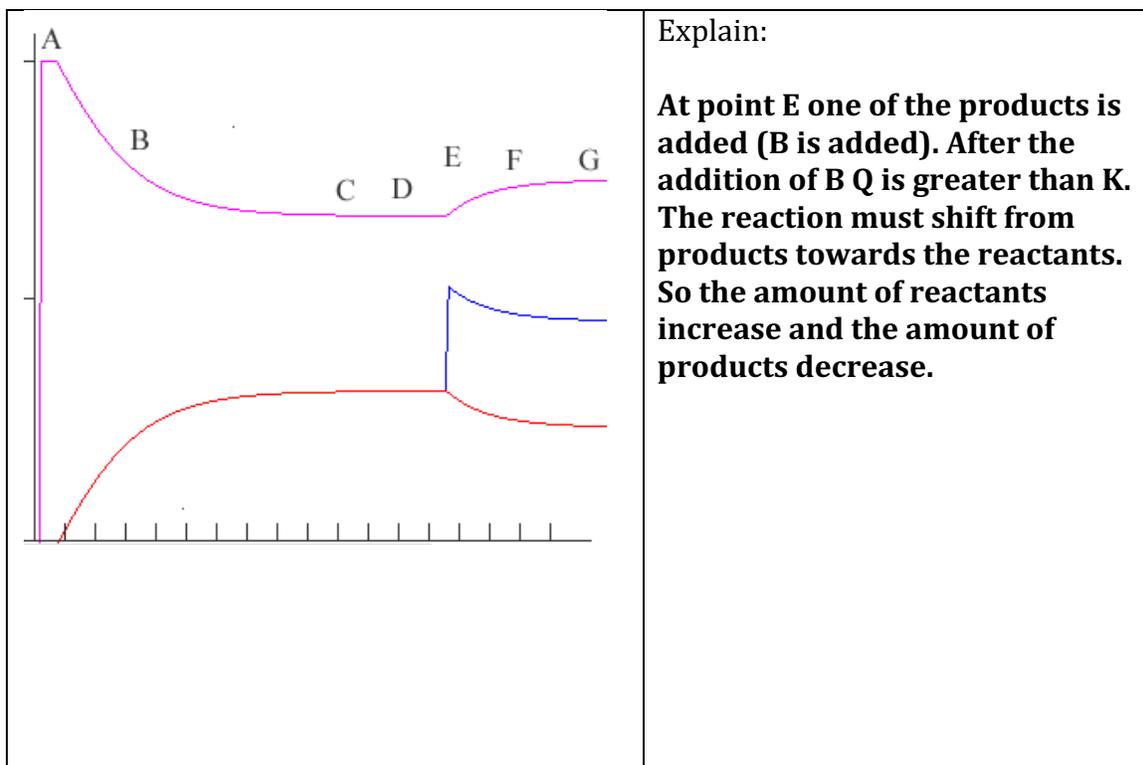
- b) Indicate whether K for the reaction is greater than 1, less than 1 or equal to 1. Explain.

$$K = \frac{[\text{B}][\text{R}]}{[\text{BR}]} = \frac{\text{small amount}}{\text{large amount}} \quad K \text{ for the reaction is less than 1.}$$

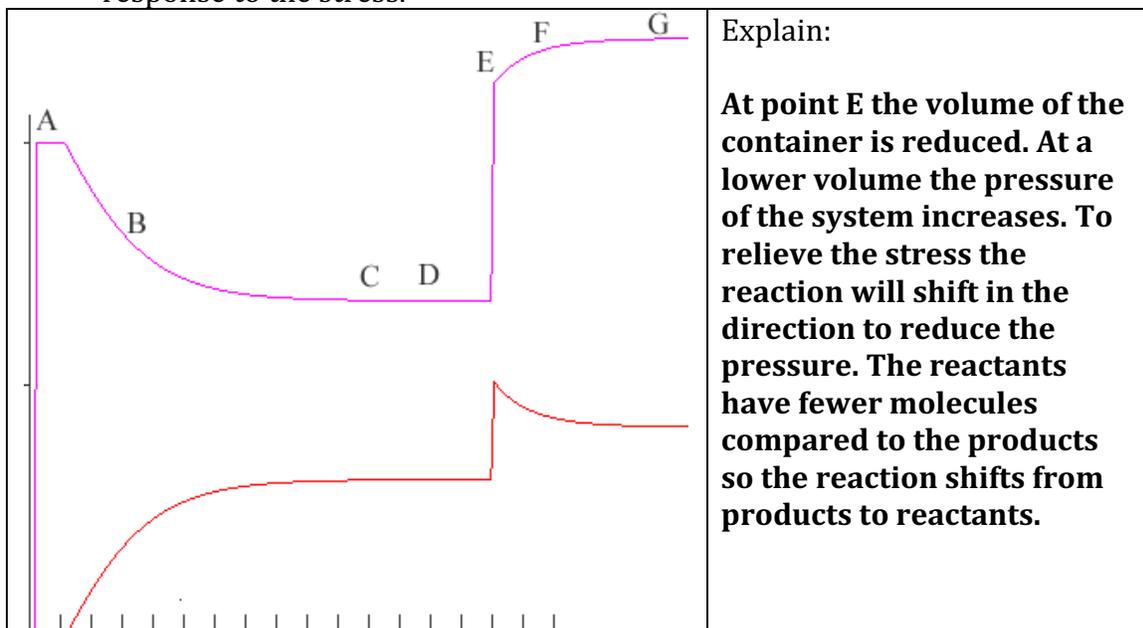
- c) At point 'B' indicate how Q compares to K . Explain.

At point B $Q < K$. At point B there is a lower amount of products and a higher amount of reactants.

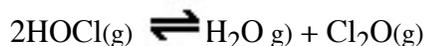
- d) In this new view the same reaction has occurred. Indicate the stress (at point E) that was imposed on the system, and explain how the system changed as a response to the stress.



- e) In this new view the same reaction has occurred. Indicate the stress (at E) that was imposed on the system, and explain how the system changed as a response to the stress.



4. a) The following reversible reaction occurs at 25. °C.



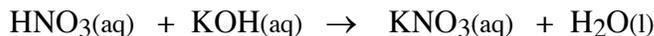
A reaction vessel that contains 0.500 atm of HOCl initially is allowed to attain equilibrium at 25 °C. Analysis indicates 0.0650 atm of HOCl are present at equilibrium. Calculate K_p for the reaction as written.

$2\text{HOCl(g)} \rightleftharpoons \text{H}_2\text{O(g)} + \text{Cl}_2\text{O(g)}$	$K_p = \frac{P(\text{H}_2\text{O}) (P(\text{Cl}_2\text{O}))}{P(\text{HOCl})^2} = \frac{P(0.217) (0.217)}{P(0.0650)^2} = 11.1$
<p>Initial 0.500 atm 0 0</p> <p>Change</p> <p>Equilibrium 0.0650 atm</p>	
$2\text{HOCl(g)} \rightleftharpoons \text{H}_2\text{O(g)} + \text{Cl}_2\text{O(g)}$	
<p>Initial 0.500 atm 0 0</p> <p>Change -0.435 atm +0.217 +0.217</p> <p>Equilibrium 0.0650 atm +0.217 +0.217</p>	

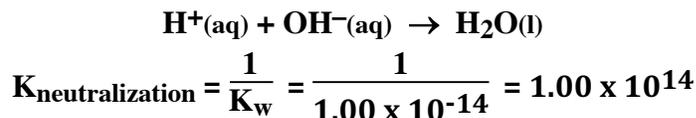
b) After the reaction in 4a attains equilibrium, some HOCl is added so the new partial pressure of HOCl is 0.200 atm. Calculate the new equilibrium amounts of the system when equilibrium is re-established.

$2\text{HOCl(g)} \rightleftharpoons \text{H}_2\text{O(g)} + \text{Cl}_2\text{O(g)}$	$P(\text{H}_2\text{O}) = P(\text{Cl}_2\text{O}) = 0.217 \text{ atm} + 0.059 \text{ atm} = 0.276 \text{ atm}$ $P(\text{HOCl}) = 0.200 \text{ atm} - 2 \cdot 0.059 \text{ atm} = 0.082 \text{ atm}$
<p>Initial 0.200 atm +0.217 +0.217</p> <p>Change -2x +x +x</p> <p>Equilibrium 0.200-2x 0.217+x 0.217+x</p>	
$11.1 = \frac{P(\text{H}_2\text{O}) P(\text{Cl}_2\text{O})}{P(\text{HOCl})^2} = \frac{(0.217+x) (0.217+x)}{(0.200-2x)^2}$	
$11.1 = \frac{(0.217+x)^2}{(0.200-2x)^2}$	
<p>Take the square root of both side</p>	
$3.34 = \frac{0.217+x}{0.200-2x}$	
$3.34 (0.200-2x) = 0.217+x$	
$0.668 - 6.68x = 0.217+x$	
$0.451 = 7.68x$	
$x = 0.059 \text{ atm}$	

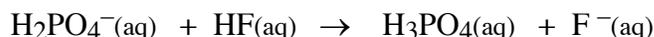
5. In the reaction.



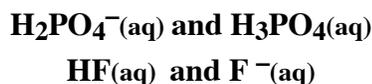
a) Write the net ionic equation and determine the value of K for the neutralization reaction between nitric acid and potassium hydroxide?



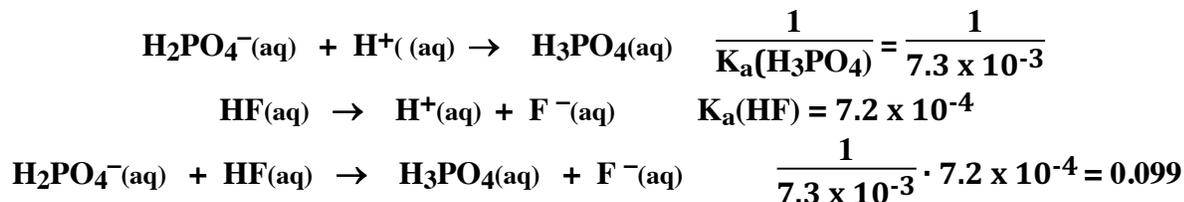
i) In the neutralization reaction between $\text{H}_2\text{PO}_4^-(\text{aq})$ and $\text{HF}(\text{aq})$



i) Identify a conjugate acid-base pair in the equation above.



ii) Determine K for the reaction above. Show your work. (see table of K's at the end of this document.)



ii) Which acid in the above reaction is the strongest? What evidence did you use to make the claim?

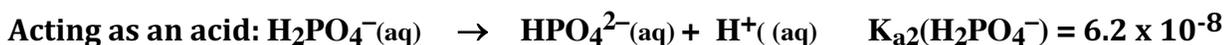
H_3PO_4 is the stronger acid. K for the reaction is small, favoring the reactants. The $[\text{HF}]$ is greater than the $[\text{H}_3\text{PO}_4]$ at equilibrium. The stronger the acid the smaller the concentration of the.

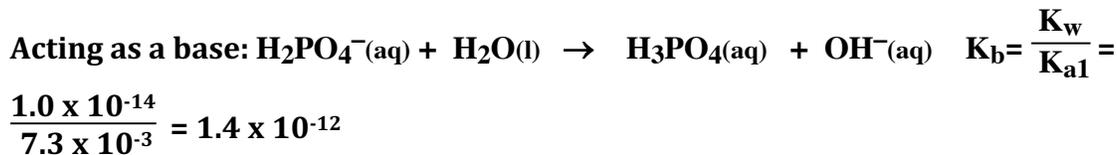
6. Phosphoric acid, $\text{H}_3\text{PO}_4(\text{aq})$ is a triprotic acid with the following equilibrium constants; $K_{a1} = 7.3 \times 10^{-3}$, $K_{a2} = 6.2 \times 10^{-8}$, and $K_{a3} = 4.2 \times 10^{-13}$.

a) Is a solution that is 0.1 M $\text{H}_3\text{PO}_4(\text{aq})$ acidic, basic or neutral? Explain.

H_3PO_4 is an acid, so an aqueous solution will be acid.

b) Is a solution that is 0.1 M $\text{NaH}_2\text{PO}_4(\text{aq})$ acidic, basic or neutral? Explain.

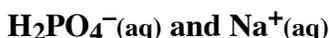




The larger K_a compared to K_b means H_2PO_4^- acts as an acid.

c) Is a solution that is 0.1 M $\text{Na}_2\text{HPO}_4(\text{aq})$ acidic, basic or neutral? Explain.
Acidic. See Part b of this question.

d) Given that NaH_2PO_4 is soluble in water identify the two ions present when NaH_2PO_4 dissolves in water.



e) Of the two ions in the NaH_2PO_4 solution, identify the ion that can act as a Brønsted-Lowry base?



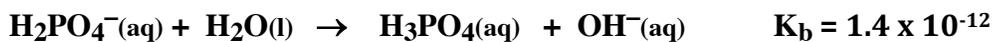
f) The weak acid, $\text{HC}_2\text{H}_3\text{O}_2$ ($K_a = 1.8 \times 10^{-5}$) is added to the solution of NaH_2PO_4 .

i) Write a Brønsted-Lowry acid-base balanced chemical equation between $\text{HC}_2\text{H}_3\text{O}_2$ and the ion in the solution that will behave as a base and label all species as acid, base, conjugate acid or conjugate base.



This reaction is composed of two reactions:

ii) The K for the above reaction has a value of 2.33×10^{-3} . Do you agree or disagree? Support your claim with an explanation.



$$K = K_a \cdot K_b \cdot \frac{1}{K_w} =$$

iii) Which of the two acids is the strongest?

Acetic acid is the strongest acid because the equilibrium constant for acetic acid (1.8×10^{-5}) is larger than the K_{a2} (6.2×10^{-8}).

7. In aqueous solution, trimethylamine, $(\text{CH}_3)_3\text{N}$, acts as a weak base. The pH of 0.250 M $(\text{CH}_3)_3\text{N}$ is 11.60
- a) Write the balanced chemical equation that describes how $(\text{CH}_3)_3\text{N}$ acts as a weak base.



- b) Write the equilibrium constant expression, K_b , for trimethylamine. (3)

$$K_b = \frac{[(\text{CH}_3)_3\text{NH}^+][\text{OH}^-]}{[(\text{CH}_3)_3\text{N}]}$$

- c) Calculate the value of the $[\text{OH}^-]$ in 0.250 M $(\text{CH}_3)_3\text{N}$.

$$\begin{aligned} \text{pOH} &= 14 - \text{pH} = 14 - 11.60 = 2.40 \\ [\text{OH}^-] &= 10^{-\text{pOH}} = 10^{-2.40} = 4.0 \times 10^{-3} \text{ M} \end{aligned}$$

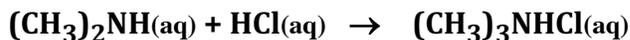
- d) Calculate the value of K_b for the reaction.

	$(\text{CH}_3)_3\text{N}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow (\text{CH}_3)_3\text{NH}^+(\text{aq}) + \text{OH}^-(\text{aq})$			
Initial	0.250 M	-	0	~0
Change	-x	-	+x	+x
Equilibrium	0.250 - x	-	+x	+x

$$\begin{aligned} K_b &= \frac{[(\text{CH}_3)_3\text{NH}^+][\text{OH}^-]}{[(\text{CH}_3)_3\text{N}]} = \frac{x \cdot x}{0.250 - x} = \frac{4.0 \times 10^{-3} \cdot 4.0 \times 10^{-3}}{0.250 - 4.0 \times 10^{-3}} \\ &= 6.4 \times 10^{-5} \end{aligned}$$

Dimethylamine, $(\text{CH}_3)_2\text{NH}$ is also a weak base. A 25.0 mL sample of 0.350 M $(\text{CH}_3)_2\text{NH}$ is titrated with 0.500 M HCl. The value of K_b for $(\text{CH}_3)_2\text{NH}$ is 5.4×10^{-4} .

- e) Calculate the volume of 0.500 M HCl that must be added to reach the equivalence point of the titration.



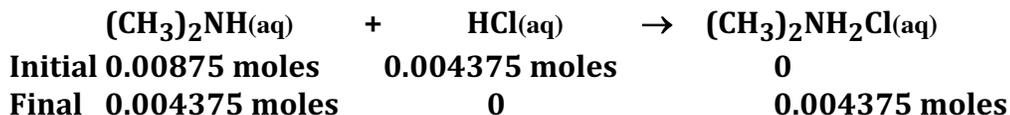
$$\text{mole}_{\text{acid}} = \text{mole}_{\text{base}}$$

$$\text{mole}_{\text{base}} = 0.025 \text{ L} \cdot 0.350 \frac{\text{mol}}{\text{L}} = 0.00875 \text{ mol}$$

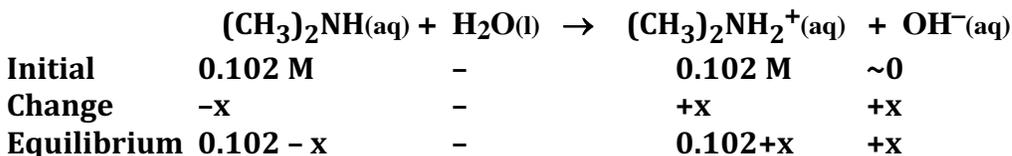
$$\text{volume}_{\text{acid}} = \frac{\text{mole}_{\text{acid}}}{\text{molarity}} = \frac{\text{mole}_{\text{acid}}}{\text{molarity}} = \frac{0.00875 \text{ mol}}{0.500 \frac{\text{mol}}{\text{L}}} = 0.0175 \text{ L}$$

- f) A student claims the $[\text{OH}^-] = K_b$ at the half-equivalence point of the titration. Provide evidence that supports the claim.

In this example the half-equivalence point occurs when exactly half of the moles of base have reacted. This will occur when 0.001875 moles of HCl have been added.



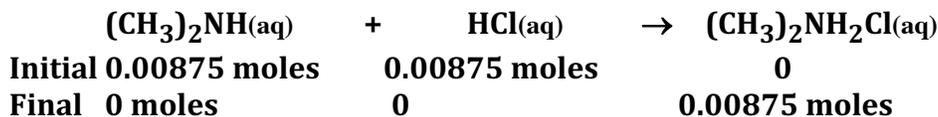
At the half-equivalence point, 17.5 mL of 0.500 M HCl was added, and

$$[(\text{CH}_3)_2\text{NH}] = \frac{0.004375 \text{ moles}}{0.0425 \text{ L}} \text{ and } [(\text{CH}_3)_2\text{NH}_3\text{NH}^+] = \frac{0.004375 \text{ moles}}{0.0425 \text{ L}} = 0.102 \text{ M}$$


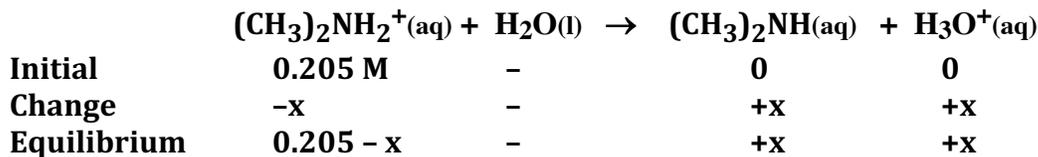
$$K_b = \frac{[(\text{CH}_3)_2\text{NH}_2^+][\text{OH}^-]}{[(\text{CH}_3)_2\text{NH}]} = \frac{0.102+x \cdot x}{0.102-x} = \frac{0.102 \cdot [\text{OH}^-]}{0.102} = [\text{OH}^-]$$

- g) The pH at the equivalence point of the titration is 5.71. Show the calculation that produces this pH.

At the equivalence point the moles of $(\text{CH}_3)_2\text{NH}$ have completely reacted with the moles of HCl, so only $(\text{CH}_3)_2\text{NH}_2^+$ remains in solution.



$$[(\text{CH}_3)_2\text{NH}_2^+] = \frac{0.00875 \text{ mol}}{0.0425 \text{ L}} = 0.205 \text{ M}$$



$$\frac{K_w}{K_b} = K_a = \frac{[(\text{CH}_3)_2\text{NH}][\text{H}_3\text{O}^+]}{[(\text{CH}_3)_2\text{NH}_2^+]} = \frac{x \cdot x}{0.205 - x}$$

$$\frac{1.0 \times 10^{-14}}{5.4 \times 10^{-4}} = 1.84 \times 10^{-11} = \frac{x \cdot x}{0.205}$$

$$x = [\text{H}_3\text{O}^+] = 1.9 \times 10^{-6} \text{ M}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(1.9 \times 10^{-6}) = 5.71$$

- h) Which of the two bases, trimethylamine or dimethylamine is the stronger base? Explain.

The strength of the base depends on the magnitude of K_b 's. The K_b for $(\text{CH}_3)_3\text{N}$ is 6.4×10^{-5} and for $(\text{CH}_3)_2\text{NH}$ is 5.4×10^{-4} . The larger the K_b the stronger the base.

8. Calculate the pH of each of the following,

a) 0.300 M HC₃H₅O₂ (propionic acid)

	HC₃H₅O₂(aq) + H₂O(l) → HC₃H₅O₂⁻(aq) + H₃O⁺(aq)			
Initial	0.300 M	-	0	0
Change	-x	-	+x	+x
Equilibrium	0.300 - x	-	+x	+x

$$K_a = \frac{[\text{HC}_3\text{H}_5\text{O}_2^-][\text{H}_3\text{O}^+]}{[\text{HC}_3\text{H}_5\text{O}_2]} = \frac{x \cdot x}{0.300 - x}$$

$$1.3 \times 10^{-5} = \frac{x \cdot x}{0.300}$$

$$x = [\text{H}_3\text{O}^+] = 2.0 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(2.0 \times 10^{-3}) = 2.70$$

b) 0.400 M CH₃NH₃NO₃ (CH₃NH₂ K_b 4.4 × 10⁻⁴)

	CH₃NH₃⁺(aq) + H₂O(l) → CH₃NH₂(aq) + H₃O⁺(aq)			
Initial	0.400 M	-	0	0
Change	-x	-	+x	+x
Equilibrium	0.400 - x	-	+x	+x

$$\frac{K_w}{K_b} = K_a = \frac{[\text{CH}_3\text{NH}_2][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{NH}_3^+]} = \frac{x \cdot x}{0.400 - x}$$

$$\frac{1.0 \times 10^{-14}}{4.4 \times 10^{-4}} = 2.3 \times 10^{-11} = \frac{x \cdot x}{0.400}$$

$$x = [\text{H}_3\text{O}^+] = 3.0 \times 10^{-6} \text{ M}$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log(3.0 \times 10^{-6}) = 5.52$$

c) 250. mLs of a solution that is 0.450 M CH₃NH₂ and 0.400 M CH₃NH₃NO₃.

	CH₃NH₂(aq) + H₂O(l) → CH₃NH₃⁺(aq) + OH⁻(aq)			
Initial	0.450 M	-	0.400 M	~0
Change	-x	-	+x	+x
Equilibrium	0.450 - x	-	0.400+x	+x

$$K_b = 4.4 \times 10^{-4} = \frac{[\text{CH}_3\text{NH}_3^+][\text{OH}^-]}{[\text{CH}_3\text{NH}_2]} = \frac{0.400+x \cdot x}{0.450 - x} = \frac{0.400 \cdot [\text{OH}^-]}{0.450}$$

$$[\text{OH}^-] = 5.0 \times 10^{-4}$$

$$\text{pOH} = -\log[\text{OH}^-] = -\log(5.0 \times 10^{-4}) = 3.30 : \text{pH} = 14 - \text{pOH}, \text{pH} = 10.70$$

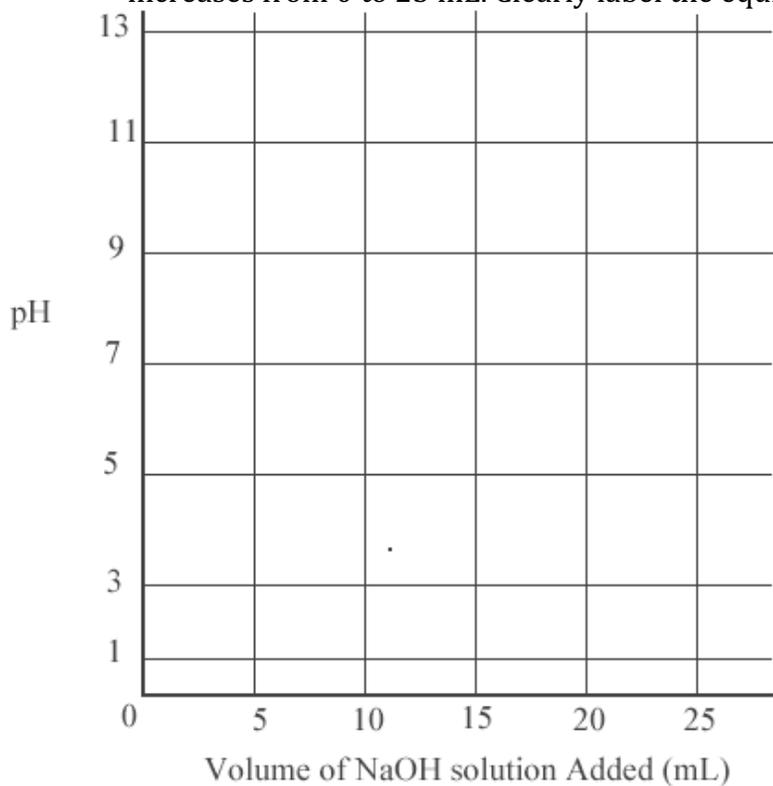
d) after adding 0.0025 mol HNO₃ to the solution in part 2c. (**Excluded**)

9. A solution of NaOH with an approximate concentration of 0.1 M is to be standardized by titration. Assume that the following materials are available.

Clean, dry 25 mL buret 125 mL Erlenmeyer flask Wash bottle filled with distilled water	Analytical balance Phenolphthalein indicator solution Potassium hydrogen phthalate, KHP ($MM = 204 \text{ g mol}^{-1}$), solid monoprotic acid (to be used as a primary standard)
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- a) Briefly describe the steps you would take, using the materials listed above, to standardize the NaOH solution. Indicate the data you must collect, and how the data is used to calculate the concentration of the NaOH solution.
- 1. Weigh 0.306 g of KHP using the analytical balance and add to the 125 mL Erlenmeyer flask. Stir the mixture until the KHP has dissolved and add a few drops of the phenolphthalein indicator to the solution.**
 - 2. Pour 5 – 10 mL of the NaOH solution into the buret, and rinse all the inside surfaces of the buret. Dispense this sample of NaOH. Fill the buret with NaOH solution to a point above the 0 mL etch. Dispense some of the solution checking there are no bubbles in the tip of the buret.**
 - 3. Record the initial volume of the NaOH solution in the buret, and carefully add the NaOH solution into the flask. Continue adding until the solution in the flask turns a light shade of pink that persists with swirling. Record the final volume of the NaOH solution in the buret.**

- b) After the NaOH is standardized (to 3 significant figures) it is used to titrate a weak monoprotic acid, HX. The equivalence point is reached after adding 15.0 mL of the NaOH solution. In the space below sketch a titration curve, showing the pH changes that occur as the volume of NaOH solution added increases from 0 to 25 mL. Clearly label the equivalence point on the curve.



- c) Describe how the value of the acid dissociation equilibrium constant, K_a , for the weak acid, HX, could be determined from the titration curve you drew in part b.

10. In aqueous solution, lactic acid, $\text{HC}_3\text{H}_5\text{O}_3$, behaves as a weak acid. In 0.0247 M $\text{HC}_3\text{H}_5\text{O}_3$ at $25\text{ }^\circ\text{C}$, the hydrogen ion concentration, $[\text{H}^+]$ is $1.79 \times 10^{-3}\text{ M}$. In answering the following questions assume that the temperature is constant at $25\text{ }^\circ\text{C}$ and that volumes are additive.

- a) Write the weak acid dissociation chemical equation for lactic acid, $\text{HC}_3\text{H}_5\text{O}_3$.

- b) write the equilibrium-constant expression for the reaction in part a).

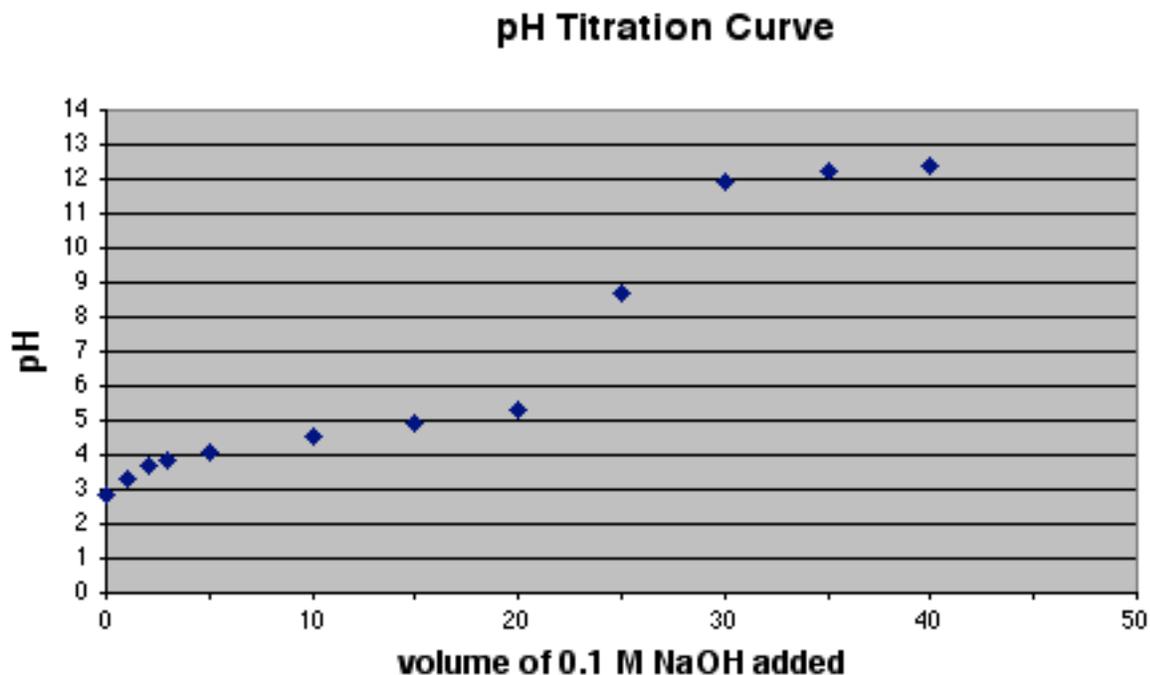
- c) Determine the pH of $0.0247\text{ M HC}_3\text{H}_5\text{O}_3(\text{aq})$.

- d) Determine the value of the acid ionization constant, K_a , for $\text{HC}_3\text{H}_5\text{O}_3(\text{aq})$.

- e) In an experiment a 20.0 mL sample of $0.0247\text{ M HC}_3\text{H}_5\text{O}_3(\text{aq})$ was placed in a flask and titrated to the equivalence point and beyond using 0.0450 M KOH .
 - i) Determine the volume of 0.0450 M KOH that was added to reach the equivalence point.

- ii) Is the pH of the mixture at the equivalence point, greater than, less than or equal to 7? Explain. (Do not do a calculation.)

11. The graph below shows the result of the titration of a sample of a 0.10 M solution of a weak acid, HA, with a strong base, 0.10 M NaOH.



- i) What are two features of the graph above that identify the acid as a weak acid?

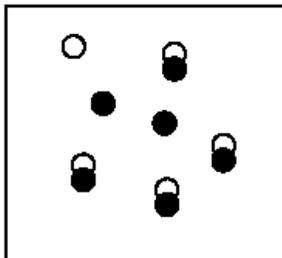
- ii) without doing calculations, describe one method by which the acid ionization constant for the unknown substance can be determined.

- iii) Using the graph above, sketch the titration curve the would be obtained if an equal amount of 0.10 M HCl were used instead of the 0.10 M HA

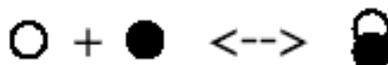
- iv) If 0.20 M NaOH were used to titrate the same volume of the 0.10 M HA, what would be the volume of base added at the equivalence point? Would the pH at the equivalence point be equal to, greater than, or less than the pH at the equivalence point using 0.10 M NaOH? Explain.

MULTIPLE CHOICE:

1. The following diagram represents a reaction chamber



where the chemical reaction,

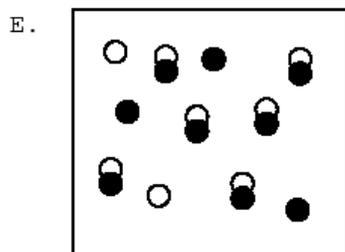
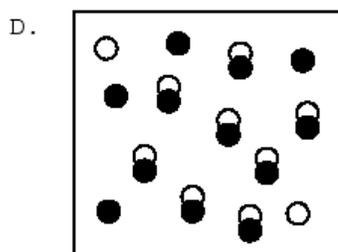
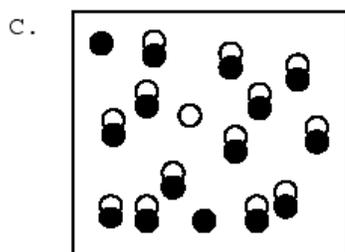
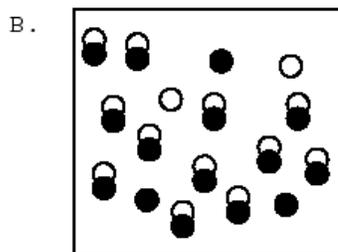
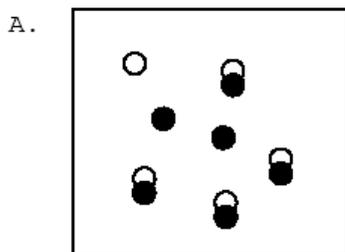


is at equilibrium

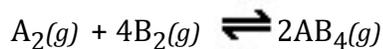
If nine units of



are added to the reaction chamber, which of the following best represents the system when it reestablishes equilibrium? **(B is correct...K value match.)**

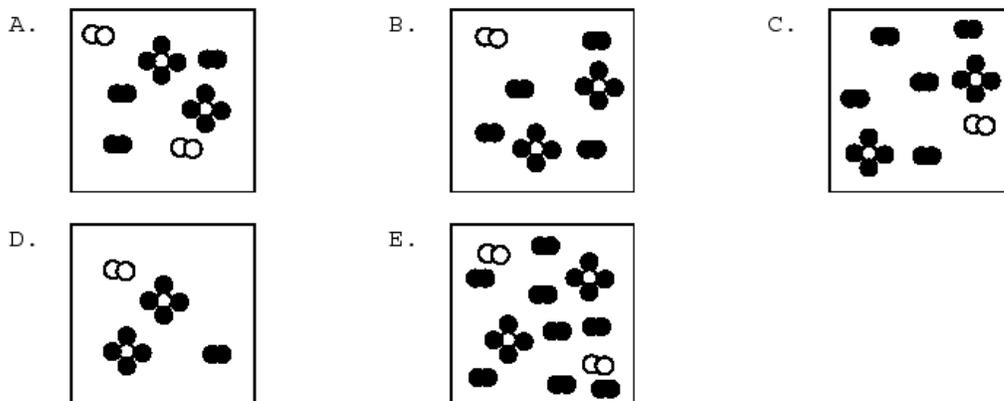


2. Consider the following hypothetical reaction:

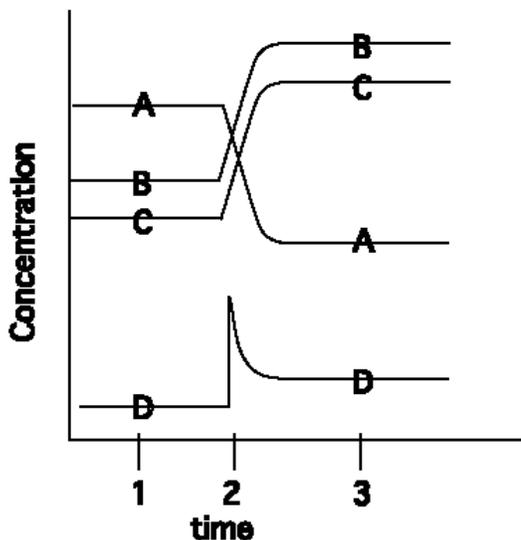


Initially 3 molecules of A_2 and 7 molecules of B_2 are placed in a reaction vessel. After a period of time 2 molecules of AB_4 are found in the vessel.

Which of the following diagrams represents this final state? **(A is correct)**



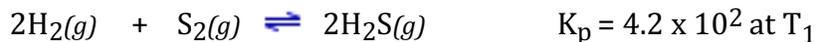
3. The following diagram represents a hypothetical chemical reaction



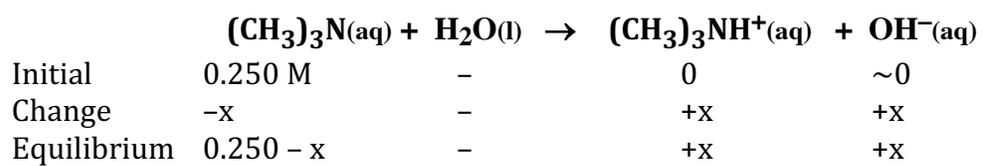
At time 2 a change takes place in the reaction. Which of the following statements about this chemical system is false? **(E is correct...D and A are on the same side of the equation).**

- A) At time 3 the system is at equilibrium.
- B) At time 2 the concentration of D was increased
- C) The change at time 2 caused more C to be formed.
- D) The diagram is an illustration of the effect that changing the concentration has on a reaction at equilibrium
- E) D and A are on opposite sides of the equation representing the reaction

Answer Questions 4 - 6 over the reaction



4. A reaction mixture contains 0.0400 atm of H_2 , 0.120 atm of S_2 and 0.560 atm of H_2S . is the reaction at equilibrium? **(B is the correct choice Q is 1633)**
- (A) No, because $Q = K_p$
(B) No, because $Q \neq K_p$
(C) Yes, because $Q = K_p$
(D) Yes, because $Q \neq K_p$
5. Which direction is the reaction proceeding based on the conditions in Q4? **(C is correct)**
- (A) both directions since the reaction is already at equilibrium since $Q = K_p$
(B) both directions since there are some amounts of all of the chemical species in the balanced chemical equation;
(C) right to left to establish equilibrium because $Q > K_p$;
(D) left to right to establish equilibrium because $Q < K_p$
6. The reaction above is exothermic. At a different temperature (T_2) K_p is 1.6×10^4 which of the following statements is correct. **(Choice D is correct)**
- (A) T_2 is greater than T_1 because the reaction is going left to right
(B) T_2 is greater than T_1 because the reaction is going right to left
(C) T_2 is less than T_1 because the reaction is going right to left
(D) T_2 is less than T_1 because the reaction is going left to right



Name	Formula	K_{a1}	K_{a2}	K_{a3}
Acetic	$\text{HC}_2\text{H}_3\text{O}_2$	1.8×10^{-5}		
Ascorbic	$\text{HC}_6\text{H}_7\text{O}_6$	8.0×10^{-3}		
Arsenic	H_3AsO_4	5.6×10^{-3}	1.0×10^{-7}	3.0×10^{-12}
Arsenous	H_3AsO_3	6.0×10^{-10}		
Benzoic	$\text{HC}_7\text{H}_5\text{O}_2$	6.5×10^{-5}		
Boric	H_3BO_3	5.8×10^{-10}		
Butyric acid	$\text{HC}_4\text{H}_7\text{O}_2$	1.5×10^{-5}		
Carbonic	H_2CO_3	4.3×10^{-7}	5.6×10^{-11}	
Cyanic	HCNO	3.5×10^{-4}		
Citric	$\text{H}_3\text{C}_6\text{H}_5\text{O}_7$	7.4×10^{-4}	1.7×10^{-5}	4.0×10^{-7}
Formic	HCHO_2	1.8×10^{-4}		
Hydroazoic	HN_3	1.9×10^{-5}		
Hydrocyanic	HCN	4.9×10^{-10}		
Hydrofluoric	HF	7.2×10^{-4}		
Hydrogen chromate ion	HCrO_4^-	3.0×10^{-7}		
Hydrogen peroxide	H_2O_2	2.4×10^{-12}		
Hydrogen selenate ion	HSeO_4^-	2.2×10^{-2}		
Hydrogen sulfate ion	HSO_4^-	1.2×10^{-2}		
Hydrogen sulfide	H_2S	5.7×10^{-8}	1.3×10^{-13}	
Hypobromous	HBrO	2.0×10^{-9}		
Hypochlorous	HClO	3.0×10^{-8}		
Hypoiodous	HIO	2.0×10^{-11}		
Iodic	HIO_3	1.7×10^{-1}		
Lactic	$\text{HC}_3\text{H}_5\text{O}_3$	1.4×10^{-4}		
Malonic	$\text{H}_2\text{C}_3\text{H}_2\text{O}_4$	1.5×10^{-3}	2.0×10^{-6}	
Oxalic	$\text{H}_2\text{C}_2\text{O}_4$	5.9×10^{-2}	6.4×10^{-5}	
Nitrous	HNO_2	4.5×10^{-4}		
Phenol	$\text{HC}_6\text{H}_5\text{O}$	1.3×10^{-10}		
Phosphoric	H_3PO_4	7.5×10^{-3}	6.2×10^{-8}	4.2×10^{-13}
Paraperiodic	H_5IO_6	2.8×10^{-2}	5.3×10^{-9}	
Propionic	$\text{HC}_3\text{H}_5\text{O}_2$	1.3×10^{-5}		
Pyrophosphoric	$\text{H}_4\text{P}_2\text{O}_7$	3.0×10^{-2}	4.4×10^{-3}	
Selenous	H_2SeO_3	2.3×10^{-3}	5.3×10^{-9}	
Sulfuric	H_2SO_4	strong acid	1.2×10^{-2}	
Sulfurous	H_2SO_3	1.7×10^{-2}	6.4×10^{-8}	
Tartaric	$\text{H}_2\text{C}_4\text{H}_4\text{O}_6$	1.0×10^{-3}	4.6×10^{-5}	

E.2 DISSOCIATION CONSTANTS FOR BASES AT 25°C

Name	Formula	K_b	Name	Formula	K_b
Ammonia	NH_3	1.8×10^{-5}	Hydroxylamine	HONH_2	1.1×10^{-8}
Aniline	$\text{C}_6\text{H}_5\text{NH}_2$	4.3×10^{-10}	Methylamine	CH_3NH_2	4.4×10^{-4}
Dimethylamine	$(\text{CH}_3)_2\text{NH}$	5.4×10^{-4}	Pyridine	$\text{C}_5\text{H}_5\text{N}$	1.7×10^{-9}
Ethylamine	$\text{C}_2\text{H}_5\text{NH}_2$	6.4×10^{-4}			
Hydrazine	H_2NNH_2	1.3×10^{-6}			