Exam 1 Chem 311 Name:
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Multiple choice (2.5 pts each for a total of 25/100 exam pts)

1. When a chemical system is at equilibrium,

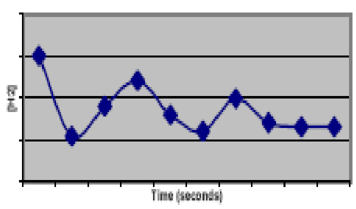
   A. the forward and reverse reactions have stopped.
   B. the reaction quotient, Q, has reached a minimum.
   C. the reaction quotient, Q, has reached a maximum.
   D. the concentrations of the reactants and products have reached constant values.
   E. the concentrations of the reactants are equal to the concentrations of the products.

2. Consider the following reaction:

$$H_2\,(g) + I_2\,(g) \leftrightarrows 2\,HI\,(g)$$

   If H_2 and I_2 are mixed together and allowed to come to equilibrium, what would the graph of the concentration of H_2 look like over time?

A



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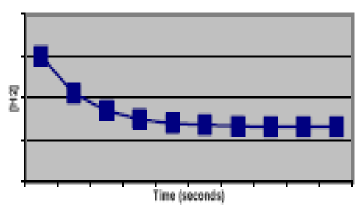

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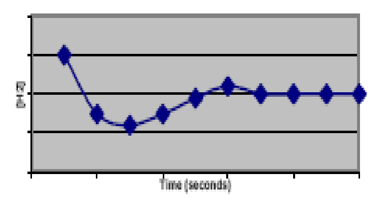

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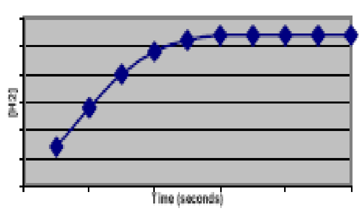

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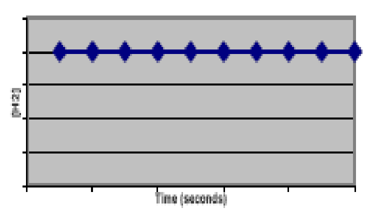


3. Consider the following reaction:

$$H_2\ (g) + F_2\ (g) \leftrightarrows 2\ HF\ (g)$$

A flask containing these three chemical species is at equilibrium. Additional H_2 is added to the flask. As the system returns to equilibrium, which of the following compounds will experience a change in concentration as the system approaches equilibrium?

A. H_2
B. F_2
C. HF
D. Two of the above
E. All of the above

4. At 25°C, the equilibrium constant *K*c for the reaction

$$2A(g) \rightarrow B(g) + C(g)$$

is 2.25. A mixture of 0.500 moles of B and C in a 10.0 L container is allowed to come to equilibrium. What is the equilibrium concentration of A?

A. 0.0100 *M*
B. 0.0500 *M*
C. 0.100 *M*
D. 0.0333 *M*
E. 0.0250 *M*

5. What distinguishes a weak acid from a strong acid dissolved in water?

A. All acids are strong.
B. A weak acid doesn't dissociate as much in water.
C. A weak acid has fewer protons to donate.
D. A weak acid will be more dilute.
E. A weak acid is smaller than a strong acid.

6. What is the pH of a 4.4 mM HCl solution?

A. 2.36
B. 11.64
C. 4.4
D. 9.6
E. none of these choices is correct

7. Assuming that only two chemicals (besides water) were available to you, which two chemicals could you use to prepare a buffer at pH 7?

   A. H_3PO_4 and HCl
   B. NaH_2PO_4 and HCl
   C. NaH_2PO_4 and K_2HPO_4
   D. NaOH and K_2HPO_4
   E. NaOH and H_2O

8. What is the pH of 1.0×10^{-10} M NaOH?

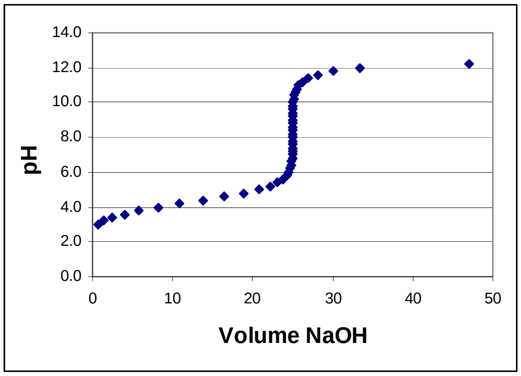
   A. 4
   B. Between 5 and 6
   C. 7
   D. Between 9 and 8
   E. 10

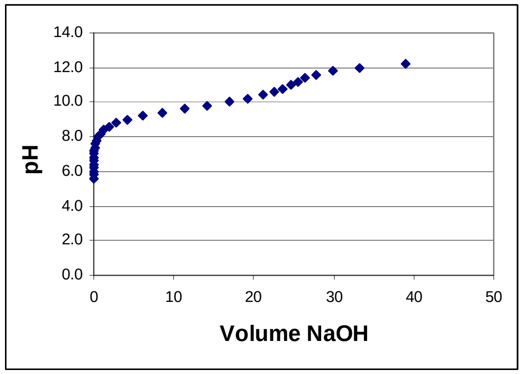
9. The solubility of silver chloride, AgCl, when placed into a 0.10 M NaCl solution __________, relative to purified water.

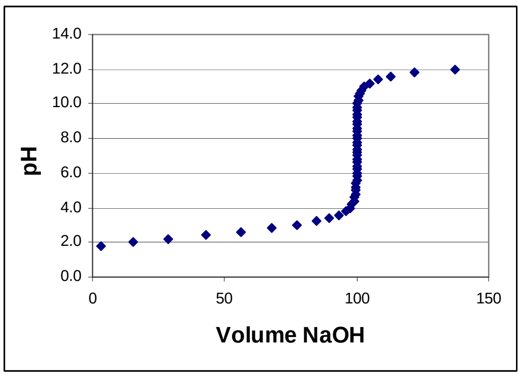
   A. decreases slightly
   B. decreases by several orders of magnitude
   C. increases slightly
   D. increases by several orders of magnitude
   E. is the same as it is in pure water

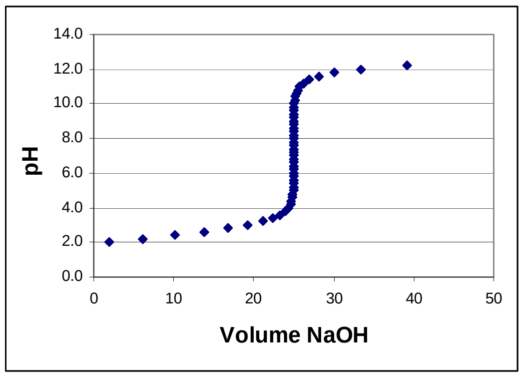
10. A 50.00 ml aliquot of a 0.0500 M solution of phenylacetic acid is titrated with 0.1000 M NaOH. Place a check mark above the titration curve that best represents this phenyacetic acid titration above.

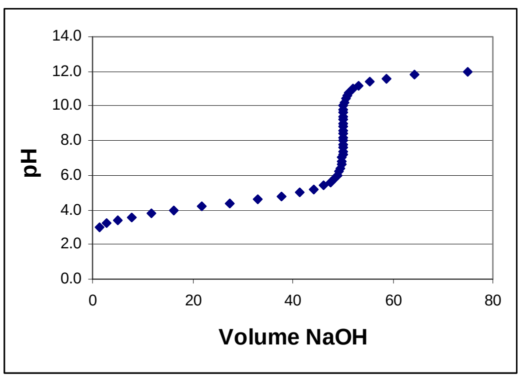
The first one (V* = 25 mL and pH =Pka = 4.3 at 12.5 mL)











Problems

1. 0.1477 g of cysteamine is transferred quantitatively to a 50 mL volumetric flask, dissolved with purified water and the flask was diluted to the 50 mL mark. 25.00 mL of this solution was quantitatively transferred to an Erlymeyer flask and titrated with 0.05000 M HCl. Predict the pH after the addition of the following volumes; 0.00 mL, 12.50 mL, 25.00 mL, 45.00 mL and 50.00 mL. (35 pts)

   Cysteamine is the most basic species in a diprotic system.
   MM(cysteamine) = 59.0907 g/mol
   $pKa_1 = 8.35$
   $pKa_2 = 10.81$

Ka1	4.47E-09
Ka2	1.55E-11
Kb1	6.52E-04
Kb2	2.26E-06

F_C = (0.1477g)*(1mol/59.0907g) / (0.05000 L) = 0.0500 M
V1* = 25.00 mL
V2* = 50.00 mL

at V = 0 mL
$Kb_1 = x^2/F_B\text{-}x$, where x = $[OH^-]$, pH = 11.76
Assumption 2 failed, but close enough for an exam

at V = 12.5 mL
You are ½ way to first equiv. pt., [B] = $[BH^+]$, so pH = pKa2 = 10.81

at V = 25.0 mL
You are at the first equivalence point, all of the original B has been converted to BH^+.
pH = (pKa1 + pKa2) / 2 = 9.58

at 45.0
You are 4/5 of the way to the second equivalence point
pH = pKa1 + log(1/4) or pH = pKa1 + log{(50-45)/(50-25)} = 7.75

at 50 mL
You are at the second equivalence point. All of the original B has been converted to BH_2^{2+}.
F_{BH22+} = 25 mL*0.0500/(25+50 mL) = 0.0167 M
Ka1 = x^2/F_{BH22+}, where $[H_3O^+]$ = x = 8.65E-6 M
pH = 5.06

2. 1.00 mL of a 0.500 M ammonia (NH_3) solution is added to 100.0 mL of a saturated solution of $Ca(OH)_2$ (obtained by adding solid $Ca(OH)_2$ salt to purified water and filtering to remove any solid suspensions). Calculate the alpha fraction of NH_4^+ and NH_3 in the resulting solution.
   HINT: External control of the pH by the $Ca(OH)_2$ (20 pts)

   $Ksp(Ca(OH)_2) = 4.68 \cdot 10^{-6}$
   $Ka\ (NH_4^+) = 5.57 \cdot 10^{-10}$, the Ka was wrong on the original test, but it did not effect how one solves the problem

$[H_3O^+]$ controlled by Ca(OH)2. So, use the Ksp equilibrium to determine $[H_3O^+]$.

$Ksp(Ca(OH)_2) = 4.68 \cdot 10^{-6} = x(2x)^2 = 4x^3$, where $x = [Ca^{2+}]$ and $2x = [OH^-]$.
x = 0.01054 M
[OH-] = 0.02107 M
$[H_3O^+]$ = 4.79E-13 M

$\alpha_{NH3} = ka / ([H_3O^+] + Ka) = 0.999$
$\alpha_{NH4+} = [H_3O^+] / ([H_3O^+] + Ka) = 8.60E-4$

3. HEPES is a common buffer used in biochemistry applications. The pKa for N2-hydroxyethylpiperazine-N'-2-ethanesulfonic acid (HEPES) is 7.56. Sigma sells dry powder HEPES packets. One packet dissolved in 1.00 L of purified water produces a 0.100 M HEPES buffer at a pH of 7.50. Determine the volume of 6.00 M HCl needed to adjust the pH of 1.00 L of buffer produced from a packet to a pH of 7.35. (20 pts)

F_{HEPES} = [HA] + [A-] = 0.100 M

First you must find [HA] and [A-] in the original 1.00 L buffer.

pH = pKa + {[A-]/[HA]} = pKa + {x/(0.100-x)}
7.50 = 7.56 + log{x/(0.1-x)}
x = 0.046_{55} M = [A-] = mol A- (1.00 L of buffer)
[HA] = 0.100-x = 0.053_{45} M

OK. Now you know the concentrations of the conjugate base and acid in the original buffer. HCl was added to adjust the pH to 7.35.

7.35 = 7.56 + log {(0.04655-x)/(0.05345+x)}
x = 0.00841 mol

(0.00841 mol H+)(1000 mL/6.00 mol) = 1.40 mL