Chemistry 102 Exam 3 – Summer 2006

This exam consists of 30 multiple choice questions. You will have 90 minutes to complete the exam. Circle the answer on your form and bubble in the correct answer on our scantron. Please write your email address on the scantron.

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Useful equations and constants

\[ K_w (25 \, ^\circ C) = 1.0 \times 10^{-14} \quad K_w = [H_3O^+][OH^-] \quad K_w = K_a \cdot K_b \]

14 = pH + pOH

\[ K = \frac{[\text{products}]}{[\text{reactants}]} \]

\[ \% \text{Ionization} = \frac{[x]}{[\text{acid}]} \times 100 \quad \text{OR} \quad \% \text{Ionization} = \frac{[x]}{[\text{base}]} \times 100 \]

**K_a values**

- **Acid** | **K_a**
  - HF | $7.2 \times 10^{-4}$
  - HNO_2 | $5.56 \times 10^{-10}$
  - HCN | $6.2 \times 10^{-10}$
  - NH_4^+ | $1.8 \times 10^{-5}$
1) Which of the following is the correct equilibrium constant expression for the unbalanced reaction?

\[
4 \mathrm{Fe(s)} + 3 \mathrm{O}_2(g) \leftrightarrow 2 \mathrm{Fe}_2\mathrm{O}_3(s)
\]

a) \( K = \frac{[\mathrm{Fe}_2\mathrm{O}_3]}{[\mathrm{Fe}][\mathrm{O}_2]} \)

b) \( K = \frac{1}{[\mathrm{O}_2]} \)

c) \( K = \frac{[\mathrm{Fe}_2\mathrm{O}_3]^2}{[\mathrm{Fe}]^4[\mathrm{O}_2]^3} \)

d) \( K = [\mathrm{O}_2]^3 \)

e) \( K = \frac{1}{[\mathrm{O}_2]^3} \)

\[
K = \frac{[\text{products}]}{[\text{reactants}]}
\]

*only aqueous and gaseous reactants and products contribute to the equilibrium constant \( K \)

2) If \( K = 3.2 \) for this reaction, what will happen to the amount of \( \text{Fe(s)} \) if the concentration of \( \text{O}_2 \) is 1.8 M?

a) amount \( \text{Fe} \) will increase

b) amount of \( \text{Fe} \) will decrease

c) amount of \( \text{Fe} \) will stay the same

d) not enough information to answer this question.

\[
K = 3.2 \quad Q = \frac{1}{(1.8)^3}
\]

\( Q < K \) so reaction shifts right + amount of \( \text{Fe} \)

3) Which of the following is not a correct description of a chemical system which is at equilibrium?

a) \( Q = K \)

b) Change in reactant concentration = 0

c) Forward reaction rate = 0 \( \frac{\text{forward reaction rate}}{\text{reverse reaction rate}} \)

d) Change in product concentration = 0

e) All of the above are correct statements.
At 50°C, the equilibrium constant for the autoionization of water (also known as the dissociation constant for water) is $5.47 \times 10^{-14}$. Use this information to answer the next two questions.

4) Is the autoionization of water an endothermic or exothermic reaction?
   a) It is endothermic because as temperature increases, $K$ increases.
   b) It is endothermic because as temperature increases, $K$ decreases.
   c) It is exothermic because as temperature increases, $K$ increases.
   d) It is exothermic because as temperature increases, $K$ decreases.
   e) It is impossible to determine if the reaction is endothermic or exothermic without more information.

5) What is the pH of a neutral solution of water at 50°C?
   a) 0.74  
   b) 6.63  
   c) 7.00  
   d) 7.37  
   e) 13.26  
   The correct answer is c) 7.00.

6) Which of the following scenarios will shift the equilibrium towards the products?
   $2 \text{ mol gas} \ [\text{heat}] + 4 \text{ mol gas} \ [\text{NH}_3(g)] \leftrightarrow \text{N}_2(g) + 3 \text{H}_2(g)$
   $\Delta H = 91.83 \text{ kJ/mol}$
   I. Add N$_2$ shift left
   II. Remove H$_2$ shift right
   III. Increase Temperature shift right
   IV. Decrease Temperature shift left
   V. Increase Volume shift right (shifts to the side with the most moles of gas)
   a) II, III, V
   b) IV
   c) II, IV, V
   d) II, V
   e) I, II, III
   The correct answer is a) II, III, V.
7) Arrange the following 0.20M solutions according to increasing acidity:

a) HF < HI < NaF < NaI
b) HI < HF < NaF < NaI
c) NaF < NaI < HF < HI
d) NaF < NaI < HI < HF
e) NaI < NaF < HI < HF

8) Calculate the [H\(^+\)] at equilibrium of a solution made by mixing 200 mL of water with 500 mL of 0.05 M HCl.

\[
\text{[H}^+\text{]} = \frac{0.025 \text{ mol}}{0.7 \text{ L}} = 0.036 \text{ M}
\]

9) Which of the following solutions (each with a concentration of 1.0 M) would have the lowest pH?

a) NaOH \(p_\text{H} \approx 7\)
b) NaCN \(p_\text{H} \approx 7\)
c) NaCl \(p_\text{H} \approx 7\)
d) NaF \(p_\text{H} \approx 7\)
e) NaNO\(_2\) \(p_\text{H} \approx 7\)

10) A 0.50 M solution of a weak acid HA has the same pH as a 0.015 M HCl solution. Calculate the \(K_a\) for HA.

\[
\text{[H}^+\text{]}_{\text{HA}} = \text{[H}^+\text{]}_{\text{H}_2\text{O}} = 0.015 \text{ M} = x
\]

\[
HA \rightleftharpoons H^+ + A^-
\]

\[
I \quad 0.5 \quad 0 \\
C \quad -x \\
E \quad 0.5-x
\]

\[
K_a = \frac{x^2}{0.5-x} \geq \frac{x^2}{0.5}
\]

\[
K_a = \frac{(0.15)^2}{0.5} = K_a
\]
Trimethyl amine \((\text{CH}_3)_3\text{N}\) is a weak base that ionizes in water as follows:

\[
\text{CH}_3\text{N}^+(\text{aq}) + \text{H}_2\text{O}(\ell) \rightleftharpoons (\text{CH}_3)_2\text{NH}^+(\text{aq}) + \text{OH}^-\text{(aq)}
\]

A 0.120 M solution of \((\text{CH}_3)_3\text{N}\) is 2.29% ionized at 25°C. Use this information to answer the next two questions.

11) Calculate the \(K_b\).

\[
K_b = \frac{x^2}{0.12-x}
\]

\[
k_b = \frac{(0.002)^2}{0.12}
\]

\[
x = (0.3 \times 10^{-5})
\]

a) \(6.3 \times 10^{-5}\) b) \(3.2 \times 10^{-3}\) c) \(6.3 \times 10^{-1}\) d) \(1.9 \times 10^1\) e) \(4.4 \times 10^1\)

12) Calculate the pH of the solution.

\[
\% \text{ ionized} = 2.29 = \frac{x}{0.12} \times 100
\]

\[
x = 0.0027 = [\text{OH}^-]
\]

\[
\text{pOH} = -\log([\text{OH}^-]) = -\log(0.0027) = 2.56
\]

\[
\text{pH} = 14.00 - 2.56 = 11.44
\]

13) Calculate the pH of a 1.00 \times 10^{-12} M solution of potassium hydroxide.

a) 2.00 b) 2.30 c) 7.00 d) 11.69 e) 12.00

\[
\text{pH} = 7.00
\]

\[
\text{[OH}^-\text{]_{H}_2\text{O}} > \text{[OH}^-\text{]_{KOH}}
\]

\[
x \times 10^{-7} \text{M} > 1 \times 10^{-12} \text{M}
\]

14) Which of the following solutions would be most resistant to changes in pH?

a) 0.05 M NaHCO\(_3\) / 0.05 M Na\(_2\)CO\(_3\)

b) 0.10 M NaHCO\(_3\) / 0.10 M Na\(_2\)CO\(_3\)

c) 0.50 M NaHCO\(_3\) / 0.50 M Na\(_2\)CO\(_3\)

d) 1.0 M NaHCO\(_3\) / 1.0 M Na\(_2\)CO\(_3\)

e) All equally resistant to changes in pH

*The best buffer has large and equal concentrations of a weak acid and its conjugate base*
For the next four questions, calculate the pH of 20.0 mL of each of the following substances.

15) 0.200 M \( \text{C}_2\text{H}_5\text{NH}_2 \text{H}^+ + \text{Cl}^- \)

\( \text{a} \) 0.699  \( \text{b} \) 5.72  \( \text{c} \) 7.00  \( \text{d} \) 8.28  \( \text{e} \) 13.3

\[ \text{C}_2\text{H}_5\text{NH}_2 \text{H}^+ = \text{H}^+ + \text{C}_2\text{H}_5\text{NH}_2 \text{N}^+ \]

\[ K_a = 1.79 \times 10^{-11} = \frac{x^2}{0.2} \]

\[ E_{0.2-x} + x + x \]

\[ x = [\text{H}^+] = 1.89 \times 10^{-6} \text{ M} \]

\[ \text{pH} = -\log(1.89 \times 10^{-6}) = 5.72 \]

16) 0.200M \( \text{HNO}_3 \)

\( \text{c} \) 0.699  \( \text{b} \) 3.45  \( \text{c} \) 7.00  \( \text{d} \) 13.3  \( \text{e} \) none of the above

\[ [\text{H}^+] = 0.20 \text{ M} \]

\[ \text{pH} = -\log(0.2) = 0.699 \]

17) 0.200M \( \text{Ca(OH)}_2 \)

\( \text{a} \) 0.398  \( \text{b} \) 0.699  \( \text{c} \) 7.00  \( \text{d} \) 13.3  \( \text{e} \) 13.6

\[ [\text{OH}^-] = 2[\text{Ca(OH)}_2] = 0.4 \text{ M} \]

\[ \text{pOH} = -\log(0.4) = 0.398 \]

\[ \text{pH} = 14.00 - 0.398 = 13.6 \]

18) 0.200M \( \text{NH}_3 \) with 0.100M \( \text{HCl} \) (\( K_b \) \( \text{NH}_3 \) = 1.80 \( \times \) 10\(^{-5} \))

\( \text{a} \) 0.699  \( \text{b} \) 4.74  \( \text{c} \) 7.00  \( \text{d} \) 9.26  \( \text{e} \) 13.6

\[ \text{H}^+ + \text{NH}_3 \rightarrow \text{NH}_4^+ \]

Before: 0.002M \( \text{H}^+ \)

Change: -0.002

After: 0.000

Buffer

\[ [\text{NH}_3] = [\text{NH}_4^+] \]

So \[ \text{pH} = \text{pK}_b \]

\[ \text{pH} = -\log \left( \frac{1 \times 10^{-9}}{1.8 \times 10^{-5}} \right) = 9.26 \]
19) Pick the words that best complete the following statement:

As the hydroxide ion concentration of a solution increases, pOH _______ and the [H⁺] _______.

a) increases, decreases
b) increases, increases
c) decreases, increases
d) decreases, decreases
e) none of the above

↑[OH⁻] more basic
so pH ↑ and pOH ↓
and [H⁺] ↓

20) Name the conjugate acid and conjugate base for HCO₃⁻

a) acid- H₂CO₃, base - CO₃²⁻
b) acid - CO₃²⁻, base - H₂CO₃
c) acid - H₂CO₃, base – CaCO₃
d) both a and c

e) none of the above

21) Which of the following choices correctly lists the major species of a 0.500 M solution of NH₄Cl before any reactions take place?

a) NH₄⁺, Cl⁻
b) NH₄⁺, Cl⁻, H₂O
c) NH₄Cl, H₂O
d) NH₄⁺, Cl⁻, NH₃, Cl⁻
e) NH₃, OH⁻, NH₄⁺, H₂O

22) How much 1.0 M Ca(OH)₂ is required to completely neutralize 400 mL of 0.5 M HCl?

a. 50 mL
b. 100 mL

c. 200 mL
d. 800 mL
e. none of the above

\[
\text{mol H}^+ = \text{mol OH}^-
\]
\[
V_{H^+} \times M_{H^+} = V_{OH^-} - M_{OH^-}
\]
(400 mL) 0.5 M = V_{OH^-} - 2 M

\[
V_{OH^-} = 0.5 \times 400 \text{ mL} = 200 \text{ mL}
\]

\[
\frac{V_{OH^-}}{2 \text{ M}} = 0.5 \text{ M} (400 \text{ mL}) = 100 \text{ mL}
\]
For the next three questions, consider the titration curve drawn below for the titration of 50.0 mL of a weak acid 0.200 M HA ($K_a = 4.5 \times 10^{-5}$) with 0.100 M NaOH.

\begin{align*}
\text{pH} & \quad \text{buffer region} \\
\text{ml base} & \quad 1 \quad 2 \quad 3 \quad 4
\end{align*}

23) At what points in the curve would the pH of the solution be expected to be equal to 4.34?

a) Point 1  
b) Point 2  
c) Point 3  
d) Point 4  
e) None of the above

\[ pK_a = -\log \left(4.5 \times 10^{-5}\right) = 4.34 \]

24) What volume of 0.100 M NaOH is needed to reach the equivalence point for the titration?

a) 50 mL  
b) 100 mL  
c) 150 mL  
d) 200 mL  
e) None of the above

\[
\text{mol acid} = \text{mol base}
\]

\[
\frac{\text{M}_{\text{acid}} \times V_{\text{acid}}}{\text{M}_{\text{base}} \times V_{\text{base}}} = \frac{0.2M \times 50mL}{0.1M \times V_{\text{base}}}
\]

\[
V_{\text{base}} = \frac{(50mL \times 0.2M)}{0.1M} = 100mL
\]

25) In what region of the curve will the predominate species be a buffer solution?

a) Between point 1 and point 2  
b) Between point 1 and point 3  
c) Between point 2 and point 3  
d) Between point 2 and point 4  
e) None of the above
26) Which of the following buffer solutions will exhibit the greatest pH?

a) 0.100 M NH₃ with 0.100 M NH₄⁺

b) 0.200 M NH₃ with 0.100 M NH₄⁺

c) 0.400 M NH₃ with 0.400 M NH₄⁺

d) 0.800 M NH₃ with 0.200 M NH₄⁺

e) 1.00 M NH₃ with 1.00 M NH₄⁺

₂H > ₚKₐ

When

[ NH₃ ] > [ NH₄⁺ ]

27) How many of the following solutions are basic?

<table>
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<th>Weak Acid</th>
<th>Weak Base</th>
<th>Strong Acid</th>
<th>Strong Base</th>
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<tr>
<td>NH₄Cl</td>
<td>NH₃</td>
<td>Mg(OH)₂</td>
<td>NaNO₂</td>
</tr>
<tr>
<td>NH₄⁺</td>
<td>NH₃⁻</td>
<td>Mg²⁺</td>
<td>NO₃⁻</td>
</tr>
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</table>

a) 0 – none of them  b) 1  c) 2  d) 3  e) 4

28) H₂SO₄ is a diprotic acid with the following Kₐ values:

H₂SO₄ ⇌ HSO₄⁻ + H⁺  
Kₐ₁ = 1 x 10⁻⁷

HSO₄⁻ ⇌ SO₄²⁻ + H⁺  
Kₐ₂ = 1.2 x 10⁻²

Which of the following statements is true?

a) The pH of a 0.10 M H₂SO₄ solution should be greater than one (pH > 1.00).

b) HSO₄⁻ is a stronger acid than H₂SO₄.

c) HSO₄⁻ is a stronger base than SO₄²⁻.

d) H₂SO₄ is a stronger acid than SO₄²⁻.

e) None of the above statements is true.
29) Which of the following titrations will have the highest pH at equivalence?

a) 100.0 mL of 0.100 M NH₃ (K_b = 1.8 x 10⁻⁵) by 0.100 M HCl \[ \text{pH} < 7 \]

b) 100.0 mL of 0.100 M KOH by 0.100 M HCl \[ \text{pH} = 7 \]

c) 200.0 mL of 0.100 M HC₂H₃O₂ (K_a = 1.8 x 10⁻⁵) by 0.100 M NaOH \[ \text{pH} > 7 \]

\[ K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = 5.56 x 10^{-5} \]

d) 100.0 mL of 0.100 M HF (K_a = 7.2 x 10⁻⁴) by 0.100 M NaOH \[ \text{pH} > 7 \]

\[ K_a = \frac{[H^+][F^-]}{[HF]} = 7.2 x 10^{-4} \]

e) 100.0 mL of 0.100 M C₂H₅NH₂ (K_b = 5.6 x 10⁻⁴) by 0.100 M HCl \[ \text{pH} < 7 \]

The larger the K_a the stronger the base, the higher the pH.

30) How many of the following oxides will be acidic in water?

- CO₂
- Na₂O
- CaO
- SO₂

a) 0 – none of them  b) 1  c) 2  d) 3  e) 4 - all of them

Nonmetal oxides are acidic in H₂O.