1. A 1 L container initially holds 0.015 mol of H$_2$(g) and 0.020 mol of I$_2$(g) at 721K. What are the concentrations for H$_2$, I$_2$, and HI when the system reaches equilibrium?

   H$_2$ (g) + I$_2$ (g) ⇌ 2HI (g)

   \[ K_c = 1.16 \times 10^{-5} \]

   
   \[ [H_2] = \sim 0.015 \text{ M} \]
   \[ [I_2] = \sim 0.020 \text{ M} \]
   \[ [HI] = \sim 5.90 \times 10^{-5} \text{ M} \]

2. What’s the difference between the Arrhenius and Bronsted-Lowry definitions of acids and bases?

   **Arrhenius definitions:** acids increase [H$^+$] in water & bases increase [OH$^-$] in water; limited to aqueous solutions

   **Bronsted-Lowry definitions:** acids donate protons & bases accept protons; requires H atoms

3. What is the auto ionization process of water?

   H$_2$O (l) + H$_2$O (l) ⇌ H$_3$O$^+$ (aq) + OH$^-$ (aq)

   a. Write its ion-product constant expression and value at 25°C.

   \[ K_w = [H_3O^+][OH^-] = 1.01 \times 10^{-14} @ 25°C \]
Wksht 3.2 – Acid-Base
Equilibria I
Supplemental Instruction
Iowa State University

Leader: Deborah
Course: CHEM 178
Instructor: Miller
Date: 9/20/17

*For questions 1 and 2, fill out the rest of the ICE table on your own.*

1. If 0.250 mol NaI, 1.5 mol HOCl, and 0.68 mol HCl are placed in a 125 ml flask at 900 K, what is the composition of the equilibrium mixture (in terms of x)? \( K_c = 1.56 \).

\[
\text{NaI(aq) + 3 HOCl(aq) ⇌ NaIO(aq) + 3HCl}
\]

<table>
<thead>
<tr>
<th></th>
<th>NaI(aq)</th>
<th>3HOCl(aq)</th>
<th>NaIO(aq)</th>
<th>3HCl(aq)</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>E</td>
<td>0.25 - x</td>
<td>1.5 - 3x</td>
<td>x</td>
<td>0.68 + x</td>
</tr>
</tbody>
</table>

2. An evacuated vessel containing a small amount of powdered graphite is heated to 1080 K and then \( \text{CO}_2 \) is added to a pressure of 0.458 atm. Once the \( \text{CO}_2 \) is added, the system starts to produce \( \text{CO} \). After equilibrium is reached, the total pressure inside the vessel is 0.757 atm. Calculate \( K_p \).

\[
\text{CO}_2(g) + \text{C} (s) ⇌ 2 \text{CO} (g)
\]

<table>
<thead>
<tr>
<th></th>
<th>CO(_2)(g)</th>
<th>C(s)</th>
<th>2CO(g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>C</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>E</td>
<td>0.159 atm</td>
<td></td>
<td>0.598 atm</td>
</tr>
</tbody>
</table>

\( K_p = 2.25 \)

3. What are some reasons for why aqueous acid-base chemistry is important?
   - plays a role in the environment and physiologically

4. What’s the limitation of the Arrhenius acid-base definition?
   - limited to aqueous solutions
   a. In terms of \([\text{H}^+]\) and \([\text{OH}^-]\), how do you determine whether a solution is acidic, basic, or neutral?
      - Acidic: \([\text{H}^+] > [\text{OH}^-]\)
      - Basic: \([\text{H}^+] < [\text{OH}^-]\)
      - Neutral: \([\text{H}^+] = [\text{OH}^-]\)

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5. In terms of Bronsted-Lowry, what is the...
   a. conjugate base of CH₃COOH? CH₃COO⁻
   b. conjugate acid of H₂PO₄⁻? H₃PO₄

6. HBr and H₂SO₄ are examples of Bronsted-Lowry acids.
   a. Write reactions that show how they act as acids in water (H₂O).
      i. HBr (aq) + H₂O (l) → Br⁻ (aq) + H₃O⁺ (aq)
      ii. H₂SO₄ (aq) + H₂O (l) → HSO₄⁻ (aq) + H₃O⁺ (aq)
   b. Identify the conjugate acid/base pairs in the reactions.
      i. CB: Br⁻ (aq)  CA: H₃O⁺ (aq)
      ii. CB: HSO₄⁻ (aq)  CA: H₃O⁺ (aq)
   c. Write Kₐₑₒ expressions for these reactions.
      i. K(ᵢ) = ([H₃O⁺]·[Br⁻]) / [HBr]
      ii. K(ᵢi) = ([H₃O⁺]·[HSO₄⁻]) / [H₂SO₄]

7. NH₃ and CO₃²⁻ are examples of Bronsted-Lowry bases.
   a. Write reactions that show how they act as bases in water.
      i. NH₃ (aq) + H₂O (l) → NH₄⁺ (aq) + OH⁻ (aq)
      ii. CO₃²⁻ (aq) + H₂O (l) → HCO₃⁻ (aq) + OH⁻ (aq)
   b. Identify the conjugate acid/base pairs in the reactions.
      i. CB: OH⁻ (aq)  CA: NH₄⁺ (aq)
      ii. CB: OH⁻ (aq)  CA: HCO₃⁻ (aq)

8. What’s the equation to find pH? \( \text{pH} = -\log[H₃O⁺] \)
   a. What’s the pH of solution with an [H₃O⁺] of 6.8 x 10⁻³? Is it acidic, neutral, or basic?
      2.17 → very acidic
   b. What’s the concentration of H⁺ in a solution with a pH of 6.7?
      2.00 x 10⁻⁷ M
Wksht 3.3 – Acid-Base Equilibria II
Supplemental Instruction
Iowa State University

Leader: Deborah
Course: CHEM 178
Instructor: Miller
Date: 9/24/17

1. What are the conjugate bases of...
   a. HCO₃⁻ ? CO₃²⁻
   b. HBrO₂⁻ ? BrO₂⁻

2. What are the conjugate acids of...
   a. SO₄²⁻ ? HSO₄⁻
   b. H₂O ? H₃O⁺

3. List the 7 strong acids.
   HCl, HI, HBr, HClO₄, HClO₃, HNO₃, H₂SO₄

4. Write out the logarithmic equations for pH, pOH, and pKw.
   pH = -log[H⁺]
   pOH = -log[OH⁻]
   pkw = -log[H⁺][OH⁻]
   a. What is the equation for pkw in terms of pOH and pH? pkw = pH + pOH
   b. Given pH or pOH, how do you determine [H⁺] and [OH⁻]?
      [H⁺] = 10⁻ᵖʰᴴ
      [OH⁻] = 10⁻ᵖ𝐎ᴴ

5. What is the pOH of a solution with a pH of 3.67? 10.3

6. If a sample of soil has a pH of 4.7, what is the H⁺ concentration of the soil solution?
   2.00 x 10⁻⁵ M

7. What is the concentration of OH⁻ in a solution with an [H⁺] of 3.56 x 10⁻⁴ M?
   2.82 x 10⁻¹¹ M

8. Write the “Bronsted-Lowry” autoionization process of water.
   H₂O (l) + H₂O (l) ⇌ H₃O⁺(aq) + OH⁻(aq)
   a. Complete the following table.

<table>
<thead>
<tr>
<th></th>
<th>kw(50°C) = 5.48 x 10⁻¹⁴</th>
</tr>
</thead>
<tbody>
<tr>
<td>Neutral</td>
<td>pH = 6.63</td>
</tr>
<tr>
<td>Acidic</td>
<td>pH &lt; 6.63</td>
</tr>
<tr>
<td>Basic</td>
<td>pH &gt; 6.63</td>
</tr>
</tbody>
</table>
Wksht 3.4 – Acid-Base Equilibria II
Supplemental Instruction
Iowa State University
Leader: Deborah
Course: CHEM 178
Instructor: Miller
Date: 9/25/17

1. Fill in the blank.
   a. As \([H^+]\) increases, pH \textbf{decreases}.
   b. As pH increases, pOH \textbf{decreases}.
   c. The stronger an acid/base, the \textbf{weaker} its conjugate base/acid.
   d. The conjugate base of a weak acid is a \textbf{weak base}.

2. Indicate whether solutions with the following ion concentrations are neutral, acidic, or basic at 25°C.
   a. \([H^+] = 5.64 \times 10^{-8}\): \textbf{very slightly basic / almost neutral}
   b. \([OH^-] = 1.0 \times 10^{-7}\): \textbf{neutral}
   c. \([OH^-] = 3.21 \times 10^{-5}\): \textbf{basic}

3. What does the ionization of a strong substance look like versus that of a weak substance?
   \textbf{Strong substances} \Rightarrow \textbf{completely dissociate in water}
   \textbf{Weak substances} \iff \textbf{partially dissociate in water}

4. Which of the following diagrams best represents a strong acid, such as HCl, dissolved in water? Which represents a weak acid? Which represents a very weak acid? (The hydrated proton is shown as a hydronium ion. Water molecules are omitted for clarity.)

   \begin{itemize}
   \item [strong acid] \[\checkmark (b)\]
   \item [weak acid] \[\text{none}\]
   \item [very weak acid] \[\checkmark (d)\]
   \end{itemize}

5. How can one determine the strength of an acid given the \(K_a\) and \(pK_a\) values?
   \textbf{SA}: \(K_a > 1\) or \(pK_a < 0\)
   \textbf{WA}: \(K_a < 1\) or \(pK_a > 0\)
   a. \(H_3AsO_3\): \(K_a = 5.1 \times 10^{-10}\)
   \(HCN\): \(K_a = 4.9 \times 10^{-10}\)
   \(H_2O_2\): \(K_a = 2.4 \times 10^{-12}\)
   Rank these acids from strongest to weakest: \(H_3AsO_3 > HCN > H_2O_2\)
b. Calculate the $pK_a$ of hydroazoic acid which has a $K_a$ of $1.9 \times 10^{-5}$.

4.72

6. Write the ionization equation and the acid-dissociation constant expression for formic acid (HCOOH).

$$\text{HCOOH} (aq) + H_2O (l) \rightleftharpoons \text{HCOO}^- (aq) + H_3O^+ (aq)$$

$$K_a = \frac{[\text{HCOO}^-][H_3O^+]}{[\text{HCOOH}]}$$

7. A 0.056 M solution of hydrofluoric acid (HF) is prepared and its pH is found to be 2.24 at 25°C. Calculate its $K_a$ at this temperature.

$$K_a = 6.59 \times 10^{-4}$$

Review:

8. How does an increase in temperature affect $E_a$ and $k$? What about a temperature decrease?

As $T$ increases, $E_a$ and $k$ decrease. The reverse occurs when a temperature decreases.
Wksht 3.5 – Acid-Base
Equilibria III
Supplemental Instruction
Iowa State University

Leader: Deborah
Course: CHEM 178
Instructor: Miller
Date: 9/27/17

1. How can one determine the strength of a base given the $K_b$ and $pK_b$ values?
   SB: $K_b > 1$ or $pK_b < 0$
   WB: $K_b < 1$ or $pK_b > 0$
   a. $\text{NH}_2\text{OH}$: $K_b = 9.1 \times 10^{-9}$
      $\text{CH}_3\text{NH}_2$: $K_b = 4.4 \times 10^{-4}$
      $\text{PH}_3$: $K_b = 1.0 \times 10^{-14}$
      Calculate the $pK_b$s of these bases then list them from strongest to weakest:
      $\text{CH}_3\text{NH}_2 > \text{NH}_2\text{OH} > \text{PH}_3$

2. $K_w$ equals what in terms of $K_a$ and $K_b$?
   $K_w = K_a(\text{CA/A}) \times K_b(\text{B/CB})$
   a. Complete the following acid-base equilibria reaction
      $\text{IO}^- (\text{aq}) + \text{H}_2\text{O} (\text{l}) \rightleftharpoons \text{HIO} (\text{aq}) + \text{OH}^- (\text{aq})$
      The $K_a$ of the conjugate acid is $2.3 \times 10^{-11}$ at 25°C. Calculate $K_b(\text{IO}^-)$.
      $K_b(\text{IO}^-) = 4.35 \times 10^{-4}$

3. Would you expect a 1.0 M $\text{NaNO}_3(\text{aq})$ solution to be neutral, acidic, or basic at 25°C?
   neutral

4. Which of the following salts is expected to yield a basic solution when dissolved in water at 25°C? Select all that apply.
   a. $\text{NH}_4\text{Cl}$
   b. $\text{KI}$
   c. $\text{Na}_2\text{CO}_3$
   d. $\text{NaClO}_3$
   e. $\text{NaF}$

5. Calculate the pH of a 0.32 M $\text{H}_2\text{SO}_4 (\text{aq})$ solution at 298 K.
   $\text{pH} = 0.49$

Review:
6. Determine the reaction’s rate law given the following information:
   $\text{XH}_4 + 2 \text{O}_2 \rightarrow \text{XO}_2 + 2 \text{H}_2\text{O}$

<table>
<thead>
<tr>
<th>Trial</th>
<th>$[\text{XH}_4]$ (initial M)</th>
<th>$[\text{O}_2]$ (initial M)</th>
<th>Rate (M/s)</th>
</tr>
</thead>
</table>

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\[
\text{Rate} = k[XH_4]^2[O_2]
\]

a. Find the rate constant value, with units.
\[k = 57.8 \text{ M}^2\text{s}^{-1}\]
Wksht 3.6 – Acid-Base Equilibria cont.
Supplemental Instruction
Iowa State University
Leader: Deborah
Course: CHEM 178
Instructor: Miller
Date: 10/1/17

1. \( K_a(HBrO) = 2.5 \times 10^{-9} \) at 25°C. Write the chemical equilibrium and the mass-action expression for this \( K_a \) value.

\[
HBrO \ (aq) + H_2O \ (l) \rightleftharpoons H_3O^+ \ (aq) + BrO^- \ (aq)
\]

a. Write the chemical equilibrium that has the equilibrium constant \( K_b(\text{BrO}^-) \).

\[
\text{BrO}^- \ (aq) + H_2O \ (l) \rightleftharpoons HBrO \ (aq) + OH^- \ (aq)
\]

b. What is mass-action expression and the value of \( K_b(\text{BrO}^-) \) at 25°C?

\[ K_b(\text{BrO}^-) = 4.0 \times 10^{-6} \]

c. \( K_b(C_5H_5N) = 1.7 \times 10^{-9} \) at 25°C. Is \( \text{BrO}^- \) a stronger or weaker base than \( C_5H_5N \) in water?

stronger

d. Determine the value of the equilibrium constant \( K \) for the following equilibrium.

\[
HBrO \ (aq) + OH^- \ (aq) \rightleftharpoons BrO^- \ (aq) + H_2O \ (l)
\]

\[ K = 2.5 \times 10^5 \]

2. What is the pH of a 0.057 M aqueous solution of HClO_4?

\[ \text{pH} = 1.24 \]

3. Codeine (\( C_{18}H_{21}NO_3 \)) is a weak organic base. A \( 5.0 \times 10^{-3} \) M solution of codeine has a pH of 10.6. Calculate the value of \( K_b \) for this substance.

\[ K_b = 3.44 \times 10^{-5} \]

4. Calculate the pH of a 0.30 M CH_3COONa(aq) solution at 298 K. (\( K_a \) (conjugate acid) = \( 1.8 \times 10^{-5} \))

\[ \text{pH} = 9.11 \]

Review:

5. What is the half-life time for a first order reaction with a rate constant of \( 3.0 \times 10^{-2} \) s\(^{-1}\)?

\[ 23.1 \text{ s} \]
1. Write out the 2-stage dissociation process of H$_2$SO$_4$ in water.
   (1) H$_2$SO$_4$ (aq) + H$_2$O (l) $\rightarrow$ HSO$_4^-$ (aq) + H$_3$O$^+$ (aq)
   (2) HSO$_4^-$ (aq) + H$_2$O (l) $\rightleftharpoons$ SO$_4^{2-}$ (aq) + H$_3$O$^+$ (aq)

2. Write out the 2-stage dissociation process of H$_2$CrO$_4$ in water. ($K_{a1} = 1.5 \times 10^{-1}$, $K_{a2} = 3.2 \times 10^{-7}$)
   (1) H$_2$CrO$_4$ (aq) + H$_2$O (l) $\rightarrow$ HCrO$_4^-$ (aq) + H$_3$O$^+$ (aq)
   (2) HCrO$_4^-$ (aq) + H$_2$O (l) $\rightleftharpoons$ CrO$_4^{2-}$ (aq) + H$_3$O$^+$ (aq)
   a. **Correction: There is no negative charge on NaHCrO$_4$**
      Is NaHCrO$_4$ a stronger acid or base in water at 298 K?
      strong acid

3. a. Circle the protonated species.
   i. CN$^-$
   ii. HCN
   b. Circle the deprotonated species.
   i. N$_3^-$
   ii. HN$_3$
   c. What’s one way we can set the pH of a solution? using buffers

4. **Correction: the first pH should be 3.70 not 3.17**
   The $p$K$_a$ of HCNO is 3.70. If the pH of three HCNO(aq) solutions are set to 3.70, 5.82, and 7,
   which species is more abundant in each solution, HCNO(aq) or CNO$^-$(aq)?
   pH of 3.70 $\rightarrow$ equal
   pH of 5.82 $\rightarrow$ CNO$^-$
   pH of 7 $\rightarrow$ CNO$^-$

5. Calculate the pH of a 0.15 M solution of NH$_3$.
   **$K_a$(conjugate acid) = 5.6 \times 10^{-10}$**
   pH = 11.21
1. Find the fluoride ion concentration and pH of a solution that is 0.40 M in HF and 0.15 M in HCl.

\[ K_a(HF) = 6.8 \times 10^{-4} \]

\[ [F^-] = 1.81 \times 10^{-3} \text{ M} \]

2. Define a buffer solution. A solution that contains (high concentration of a) weak conjugate acid-base pair that resists drastic changes in pH when small amounts of SA or SB are added to them

   a. What are the two important characteristics of a buffer?

   - Buffer capacity & buffer pH range

   b. Consider a buffer of weak acid HA and its conjugate base A⁻. Describe what happens when a small amount of strong acid is added and when a small amount of strong base is added.

      When SA added: [HA] increases & [A⁻] decreases \( \rightarrow \) pH decreases slightly

      When SB added: [HA] decreases & [A⁻] increases \( \rightarrow \) pH increases slightly

3. What is the Henderson-Hasselbalch equation?

\[ \text{pH} = \text{pK}_a + \log \left( \frac{\text{base}}{\text{acid}} \right) \]

   a. Find the pH of a 0.25 M NH₃ / 0.18 M NH₄⁺ buffer. \( K_a(NH_4^+) = 5.6 \times 10^{-10} \)

   \[ \text{pH} = 9.39 \]

   b. What are the new concentrations and pH if the solution is diluted to 30%?

   \[ 0.075 \text{ M NH}_3 / 0.054 \text{ M NH}_4^+; \text{ pH} = 9.39 \]

4. How can 1.00 L of a 0.30 M HOAc(aq) / 0.27 M OAc⁻(aq) buffer be prepared? (glacial acetic acid is 17.4 M)

   (3 different ways. Refer to last slide of Lecture 13)