### 1. (20 pts; 2 points each) True or false: Place an X in the appropriate column

<table>
<thead>
<tr>
<th>Statement</th>
<th>True</th>
<th>False</th>
</tr>
</thead>
<tbody>
<tr>
<td>The pH at the half-equivalence point is always the average of the pKₐ at the beginning of a titration and at the equivalence point.</td>
<td></td>
<td>X</td>
</tr>
<tr>
<td>The histidine side chain is neutral at pH 4.</td>
<td>X</td>
<td></td>
</tr>
<tr>
<td>At equilibrium the rates of the forward and reverse reactions are both zero.</td>
<td>X</td>
<td></td>
</tr>
<tr>
<td>The Henderson-Hasselbach equation is not correct when [A⁻] is too large.</td>
<td>X</td>
<td></td>
</tr>
<tr>
<td>When we add equations we add equilibrium constants.</td>
<td>X</td>
<td></td>
</tr>
<tr>
<td>The buffering range for a weak acid is between the pKₐ and pKₐ + 2, but not below the pKₐ.</td>
<td>X</td>
<td></td>
</tr>
<tr>
<td>Kᵢ and Kₑ are equal to each other if there are no gases involved in the equilibrium.</td>
<td>X</td>
<td></td>
</tr>
<tr>
<td>The percent dissociation of a weak acid decreases as its concentration decreases.</td>
<td>X</td>
<td></td>
</tr>
<tr>
<td>In multiple equilibrium problems, all equilibria must be satisfied.</td>
<td>X</td>
<td></td>
</tr>
<tr>
<td>The equilibrium constant of a reverse reaction is the reciprocal of the equilibrium constant of the corresponding forward reaction.</td>
<td>X</td>
<td></td>
</tr>
</tbody>
</table>

Score for the page:___________
2. (24 pts) Chemical Equilibria

(a; 9 pts) State LeChatelier’s principle. Define the reaction quotient $Q$, and give the equivalent of LeChatelier’s principle in the language of thermodynamics.

$+$ 3 When stress is placed on an equilibrium, the equilibrium shifts to relieve the stress.

$+$ 3 $Q = \frac{[\text{products}]}{[\text{reactants}]}$ each raised to the power of its stoichiometric coefficient, not necessarily at equilibrium.

$+$ 3 $Q$ approaches $K$ as equilibrium is established.

(b; 3 pts) Write the $K_c$ expression for the following reaction:

$$\text{BaSO}_4 (s) \rightleftharpoons \text{Ba}^{2+} (aq) + \text{SO}_4^{2-} (aq)$$

$+$ 3 $K_c = \frac{[\text{Ba}^{2+}][\text{SO}_4^{2-}]}{1} = 1.1 \times 10^{-10}$

(c; 2 pts) Does the amount of dissolved barium ion depend on the amount of solid barium sulfate sitting at the bottom of the beaker? Circle one: YES  NO

(d; 6 pts) It turns out that the numerical value of $K_c$ is $1.1 \times 10^{-10}$. The small value for $K_c$ for some sulfate salts is the basis of selective precipitations in quantitative analysis, in which, for example, one can use a standardized sodium sulfate solution to measure the barium ion concentration in a test sample. Calculate the concentration of barium ion in solution if the $[\text{SO}_4^{2-}]$ is 75 mM. If the initial barium ion concentration before precipitation with NaSO$_4$ was 125 mM, what percentage of the barium ion remains in solution?

$$[\text{Ba}^{2+}] = \frac{K_c}{[\text{SO}_4^{2-}]} = \frac{1.1 \times 10^{-10}}{0.075} = 1.47 \times 10^{-9} \text{M}$$

$+$ 2 (or $1.5 \times 10^{-9}$ M)

$%$ in solution $= \frac{1.47 \times 10^{-9} \text{M}}{125 \text{mM}} \times 100\% = 1.2 \times 10^{-6} \%$ [i.e. just about none!]

Score for the page_________
(e; 4 pts) Why do we use NaSO₄ instead of sulfuric acid for the barium precipitation (other than safety)?

\[
\text{[SO}_4^{\text{2-}}\text{]} \text{ in sulfuric acid is low - it's mostly HS}_4\text{O}_4^- \text{, so the precipitation is less efficient in hot sulfuric acid.)}
\]

3. (18 pts) Biological Applications

(a; 6 pts) Draw the structure of the Lysine side chain at pH 7 and give its approximate pKₐ.

\[
\text{H}_3\text{N}^- \text{-- C} = \text{O} \text{\text{+1} for alkyli} \text{N}^+ \text{\text{+2} for amino group}
\]

(b; 12 pts) Explain why measuring the pH dependence of an enzymatic reaction is a useful probe into mechanism. Explain why, however, the experiment offers only a guide to identifying the active site residue type(s). As a concrete example, explain how a pH curve might suggest but not prove that a lysine residue is involved in catalysis.

Asp or Glu

- The pH dependence reflects mechanistic steps in which a proton must be abstracted or donated. The pH dependence tells us the pKₐ's of the residues that are being titrated as the reaction fails.

- The pKₐ's in active sites, however, can be wildly different from the pKₐ's of free amino acid side chains, so one cannot simply say that a pKₐ of 4 means aspartate.

pH curve: let's imagine but a reaction requires a deprotonated Asp or residue to act as a base. The curve to the left would result. But it could also be a histidine with an altered pKₐ that needs to be protonated.

Score for the page
4. (38 pts) Acid-base chemistry and multiple equilibria

Consider the diprotic acid malonic acid, with \( pK_a \)'s 2.83 and 5.89.
(Malonic acid is an inhibitor of the important metabolic enzyme succinate dehydrogenase.)

(a; 8 pts) The first \( pK_a \) is very low relative to a typical carboxylic acid like acetic acid, \( pK_a = 4.75 \). This means that malonic acid is a (circle one): stronger acid or a weaker acid than acetic acid. Sketch a structure that rationalizes this (three acceptable possible answers). On the other hand, the second \( pK_a \) of malonic acid is higher than the \( pK_a \) of acetic acid. Explain why.

(b; 4 pts) Why can we ignore the second \( pK_a \) when calculating the pH of a malonic acid solution?

(c; 6 pts) Calculate the pH of 0.150 M malonic acid, assuming (dubiously) that \( \chi \ll 0.150 \) M.
(d; 9 pts) Sketch a curve of the pH as we titrate 0.150 M malonic acid with NaOH, ignoring dilution.

Specify the pH for the beginning (the number you just calculated in c), the two half-equivalence points, and the two equivalence points (the second one is at pH = 9.53).
(e; 4 pts) We did not learn to calculate the exact pH for diprotic acid titrations, but we do know how to set up the problem. For any point on a titration of malonic acid $H_2M$ being converted to $HM^-$ and $M^{2-}$ with added NaOH, we know the $[Na^+]$ and we would like to calculate the concentrations of five other chemical species. One is hydroxide. Write down the other four. (Not a trick question)

$$[H^+], [H_2M], [HM^-], [M^{2-}]$$

(f; 7 pts) We need five equations to solve the problem. Two of them are the two $K_a$ relationships for the first and second acid dissociation equilibria. Write down the last equilibrium relationship needed and the two conservation equations that we need to have five equations. Hint: the total charge of the solution is zero.

$$K_w = [H^+][OH^-]$$

$$C_0 = [H_2M] + [HM^-] + [M^{2-}]$$

constant total $[\text{total}]$ of malonate

$$[Na^+] + [H^+] = [HM^-] + 2[M^{2-}] + [OH^-]$$

conservation of total charge (solution must be neutral)