Part I. Multiple-Choice

1. C Equilibrium constant and reaction quotient are the product of the products divided by the product of the reactants, each raised to the power that is their stoichiometric coefficient. All chemicals in the reaction are in the gas phase, so all of them are in the expression.

2. B The reaction is exothermic (heat is on the product side). If the temperature is increased, the reaction will shift toward reactants (less products), so the numerator of K will decrease and the denominator will increase. K will decrease.

3. A $K_p$ and $K_c$ are related by a power of $-\Delta n$. $\Delta n = -2$ for this reaction.

4. B $NH_3$ is a weak base and reacts with water when bubbled through it.

5. B
   \[
   K_c = \left( \frac{0.0143}{1.20} \right)^2 \left[ \frac{0.0168}{1.20} \right] \left( \frac{0.2064}{1.20} \right)^3 = 1.99
   \]

6. C In a Bronsted-Lowry reaction, the acid on the reactant side donates an $H^+$ and the base on the reactant side accepts it. Only reactions 1 and 2 have this form.

7. D In reaction #2, $HSO_3^-$ is the acid on the reactant side. Its conjugate base is less by an $H^+$.

8. D The three reactions sum to the overall reaction, so $K_1K_2K_3 = K_{overall}$, so $K_3 = K_{overall} / (K_1K_2)$

9. D Begin by assuming $x$ is small and then check to see if $K_a$ is small enough to justify this after you calculate. If $x$ is small, then $[H^+] = \sqrt {CAK_a}$. Using pH, $[H^+] = 10^{-pH}$. Solve for $K_a$, and indeed it is small enough to justify the assumption.

10. B The NaOH solution is twice as concentrated as the weak acid solution, so the half-equivalence point will occur when half as much NaOH solution has been added as the amount of weak acid solution in the beginning. That is the point where equal moles of NaOH have been added to match the moles of weak acid present originally.

11. B The titration is a weak acid with a strong base. The equivalence point will therefore be basic (above pH 7). The titration begins with the weak acid, so the starting point is an acidic value for pH (less than 7).

12. A The two acids are $H_2S$ and $HC_2H_3O_2$. The one with the larger $K_a$ is on the product side. Therefore the equilibrium would be expected to lie to the left (toward reactants).

13. C The conjugate base is when you take away an $H^+$.

14. A To get $K_b$, use the equation $K_w = K_aK_b$.

15. D Look for the conjugate acid base pair whose $pK_a$ is near the desired buffer pH.

16. C Calculate Q to find that it is equal to $3/8$ or 0.375. This is smaller than K, so the reaction proceeds forward to get to equilibrium.

17. A All of these are ionic salts. You have to look at the ions they form in water and figure out if they are weak acids or weak bases. $NH_4^+$ is a weak acid so it would make an acidic solution.

18. A P has a higher electrnegativity than As, so P draws electrons away from the two acidic hydrogens more than As does, thereby making those hydrogens in $H_3PO_3$ more acidic than the corresponding hydrogens in $H_3AsO_3$. 
Part II. Problems

Problem 1
(This was identical to Practice Problem 5)

a) NaOH is a strong base and dissociates completely.
   Therefore, \([\text{OH}^-] = 0.0010 \text{ M}\)
   \[\text{pOH} = -\log(0.0010) = 3.00\]
   \[\text{pH} = 14 - \text{pOH} = 11.00\]
   2 sig figs in the concentration means 2 sig figs are needed in the answer
   sig figs in a number that is a logarithm are counted after the decimal place because
   the numbers to the left of the decimal tell about the power of 10

b) HCl is a strong acid, aniline is a weak base. In this problem, the amount of acid that has been added is
   equal to half the amount of base initially present. Therefore, this is the half-equivalence point.
   At the half-equivalence point, \(pH = pK_a\), and \(pOH = pK_b\).
   \[\text{pOH} = pK_b = -\log K_b = -\log(4.3 \times 10^{-10}) = 9.37\]
   \[\text{pH} = 14 - \text{pOH} = 4.63\]
   this number also has 2 sig figs, for the same reason as the explanation in part (a)

c) NaHC\(_2\)O\(_4\) is an ionic salt that contains the ion H\(_2\)C\(_2\)O\(_4\)\(^-\). K\(_2\)C\(_2\)O\(_4\) is an ionic salt that contains the ion
   C\(_2\)O\(_4\)\(_2\)^-\.
   Since HC\(_2\)O\(_4\)\(^-\) and C\(_2\)O\(_4\)\(_2\)^- are conjugates of each other, and since they are added in
   approximately equal molar quantities, then this is a buffer solution.
   \[\text{[HC}_2\text{O}_4^-] = 0.014 \text{ mol} / 0.750 \text{ L}\]
   \[\text{[C}_2\text{O}_4^{2-}] = 0.010 \text{ mol} / 0.750 \text{ L}\]
   \[K_a = \frac{[C}_2\text{O}_4^{2-}][H^+]}{[HC}_2\text{O}_4^-] \quad \therefore \quad [H^+] = K_a \frac{[HC}_2\text{O}_4^-}{[C}_2\text{O}_4^{2-}]\]
   \[[H^+] = (6.40 \times 10^{-5})(0.014/0.010) = 8.96 \times 10^{-5}\]
   \[\text{pH} = -\log(8.96 \times 10^{-5}) = 4.05\]
   the answer also must have 2 sig figs because the numbers of moles were given with 2 sig figs

Problem 2
(This was identical to Practice Problem 10)

a) NaC\(_2\)H\(_3\)O\(_2\) is most basic
   because this is the only salt of the three which has an ion that is a weak base (C\(_2\)H\(_3\)O\(_2\)^-) or
   because C\(_2\)H\(_3\)O\(_2\)^- hydrolyzes in water to produce OH^- ions

b) NH\(_4\)Br is the solution that could be used as the titrant when a weak base is the analyte
   because this is the only salt of the three which has an ion that is a weak acid (NH\(_4^+\)) or
   because NH\(_4^+\) hydrolyzes with water to produce H^+ ions so this solution is an acid solution

c) KNO\(_3\) is the solution that has pH 7
   because K^+ is not an acid and NO\(_3^-\) is the conjugate of a strong acid so it won’t be a weak base in solution
Problem 3
(This was a mixture between Practice Problems 1, 7 and 8)

a) \[ K_c = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2 \cdot [\text{O}_2]} \]
\[ \left( \frac{5.00 \text{mol}}{2.00 \text{L}} \right)^2 \]

b) \[ Q = \left( \frac{3.00 \text{mol}}{2.00 \text{L}} \right)^2 \left( \frac{4.00 \text{mol}}{2.00 \text{L}} \right) = 1.39 \] (answer should have 3 sig figs)

Extra credit
The reaction is exothermic because decreasing the temperature is removing heat, and when heat is removed from an exothermic reaction, Le Chatelier’s principle says that it causes the equilibrium to shift toward products to replace some of the lost heat.