Part I. Multiple-Choice

1.	С	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.	В	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.	А	K_p and K_c are related by a power of - Δn . $\Delta n = -2$ for this reaction.					
4.	В	NH_3 is a weak base and reacts with water when bubbled through it.					
5.	В	$K_{c} = \left(\frac{0.0143}{1.20}\right)^{2} / \left[\left(\frac{0.0168}{1.20}\right) \left(\frac{0.2064}{1.20}\right)^{3} \right] = 1.99$					
6.	С	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 $[H^+] = sqrt \{C_A K_a\}$. Using pH, $[H^+] = 10^{-pH}$. Solve for K_a , and indeed it is small enough to justify the assumption.					
10.	В	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.	В	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.	А	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.	С	The conjugate base is when you take away an H ⁺ .					
14.	А	To get K_b , use the equation $K_w = K_a K_b$.					
15.	D	Look for the conjugate acid base pair whose pK _a is near the desired buffer pH.					
16.	С	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, $[OH^-] = 0.0010 \text{ M}$ pOH = -log(0.0010) = 3.00pH = 14 - pOH = 11.002 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 $pOH = pK_b$. $pOH = pK_b = -log K_b = -log(4.3 \times 10^{-10}) = 9.37$ pH = 14 - pOH = 4.63this number also has 2 sig figs, for the same reason as the explanation in part (a)
- c) NaHC₂O₄ is an ionic salt that contains the ion HC₂O₄. K₂C₂O₄ is an ionic salt that contains the ion $C_2O_4^{2-}$. Since $HC_2O_4^{-}$ and $C_2O_4^{-2-}$ are conjugates of each other, and since they are added in approximately equal molar quantities, then this is a buffer solution. $[HC_2O_4] = 0.014 \text{ mol} / 0.750 \text{ L}$ $[C_2O_4^{2^-}] = 0.010 \text{ mol} / 0.750 \text{ L}$ $[C \cap 2^{-1}]$ $[HC_0,-]$

$$K_a = \frac{[C_2O_4] [(H^-)]}{[HC_2O_4]} \therefore [H^+] = K_a \frac{[HC_2O_4] [C_2O_4]}{[C_2O_4]^2}$$

[H⁺] = (6.40 x 10⁻⁵) (0.014/0.010) = 8.96 x 10⁻⁵
pH = -log(8.96 x 10⁻⁵) = 4.05
the answer also must have 2 sig figs because the numbers of moles were given with 2 sig

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Problem 2

(This was identical to Practice Problem 10)

a) $NaC_2H_3O_2$ is most basic

because this is the only salt of the three which has an ion that is a weak base $(C_2H_3O_2)$ or because C₂H₃O₂⁻ hydrolyzes in water to produce OH⁻ ions

b) NH₄Br 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) KNO₃ is the solution that has pH 7

because K^+ is not an acid and NO₃ 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)
$$K_c = \frac{[SO_3]^2}{[SO_2]^2 [O_2]}$$

b) $Q = \frac{\left(\frac{5.00mol}{2.00L}\right)^2}{\left(\frac{3.00mol}{2.00L}\right)^2 \left(\frac{4.00mol}{2.00L}\right)} = 1.39$ (answer should have 3 sig figs)

c) Set up an ICE chart. Given are the initial concentrations of all three chemicals and the equilibrium concentration of O_2 . Since the equilibrium value of O_2 is smaller than the initial value, the reaction must proceed toward products to get to equilibrium. Thus, the change in $O_2 = x$ can be determined by subtraction: 2.0 - 1.6 = 0.4 M.

	2 SO ₂	+	O ₂	ŧ	2 SO ₃
Initial	1.5 M		2.0 M		2.5 M
Change	-2x		- <i>X</i>		+2 <i>x</i>
Equilibrium	1.5 - 2 <i>x</i>		1.6 M		2.5 + 2 <i>x</i>
	= 1.5 - 2(0.4)				= 2.5 + 2(0.4)
	= 0.7 M				= 3.3 M

Using the equilibrium values, calculate K_c

$$K_c = \frac{[3.3]^2}{[0.7]^2 [1.6]} = 13.9$$

the answer should have 3 sig figs because all of the numbers used to calculate it had 3 sig figs (I used 2 sig figs in the table to make it easier to see what the subtractions were)

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.