Answers to Practice Exam 2—Chemistry 122

[NOTE: Questions 4 and 9 do not “directly” apply to your (after Spring 2006) Exam 2; you had that kind of material on your Exam 1 (although the ideas in these problems are applied in solubility equilibria (Q vs K used to decide if a precipitate forms, Le Chatelier’s Principle can be applied to solubility equilibria as well). Also, there is no thermo on this practice Exam 2, but it will be on your Exam 2. See separate practice document.]

1(a) v
(b) i
(c) ii
(d) ii, i (two-part question)
(e) ii
(f) iv
(g) ii
(h) iii

2(a) T
(b) T
(c) F; weaker
(d) F; reaction occurs in the forward AND reverse directions AT EQUAL RATES.
(e) F; change “sum” to “product”

3(a) The point (~29.5 mL, {pH =}~10.4) (or something close to this).
(b) ~8.3 (or so)
(c) 0.123 M (or so; depends on your answer to part (a))

4.(a) $K_c = \frac{[Cl_2(g)]^2}{[CCl_4(g)]}$ (even better if subscripts “eq” are included)
(b) No. Since K is much less than 1, the reaction is “reactant favored”, and so one expects more reactants than products (roughly speaking) when equilibrium is established (starting from only reactants or only products). That means that the majority of reactants WON’T be converted to products.
(c) Forward reaction will occur. P(total) is halved when V is doubled. LeChatlier says that system will shift to alleviate the stress, so system will shift in direction to make more gases (make P(total) bigger). In this case, that is in forward direction, because 2 moles of gas are made for every 1 mole that is used (net increase in moles of gas) when forward reaction occurs.
(d) Let $x = [CCl_4]$ that reacts as system reaches equilibrium. Then $0.013 = \frac{(2x)^2}{3.2 - x}$ (use equilibrium expression in part (a) along with algebraic expressions for [Cl$_2$] and [CCl$_4$] from stoichiometry.
(e) $x = 0.1004 \Rightarrow [Cl_2]_{eq} = 0.201 M; [CCl_4] = 3.1 M$. Consistent with (b) in that the vast majority of CCl$_4$ did not react (only about 3% was converted to products).

5.(a) ClO$^-$
(b) ClO\(^{-}\)(aq) + H\(_2\)O(l) $\leftrightarrow$ HClO(aq) + OH\(^{-}\)(aq); \hspace{1cm} K\text{\textsubscript{b}} = [HClO][OH\(^{-}\)]/[ClO\(^{-}\)] (with eq’s)

(c) [2\textsuperscript{nd} b; oops] pH = 10.39

(d) [c; oops] pH = 7.38

6. (a) H\(_2\)PO\(_4\)\(^{-}\) + HCN $\leftrightarrow$ CN\(^{-}\) (conjugate base) + H\(_3\)PO\(_4\) (conjugate acid)
(b) REACTANT-favored. H\(_2\)PO\(_4\)\(^{-}\) competes with CN\(^{-}\) for the H\(^{+}\). If H\(_2\)PO\(_4\)\(^{-}\) is weaker, then it “loses”, and thus it remains H\(_2\)PO\(_4\) (primarily) while CN “wins”, becoming HCN (primarily) at equilibrium. Since H\(_2\)PO\(_4\)\(^{-}\) and HCN are REACTANTS in the equation as written, it must be “reactant-favored”. (Another way to answer: In the equation shown, H\(_2\)PO\(_4\)\(^{-}\) is acting as a base. But it is the WEAKER base, so the reaction as written doesn’t “go” very much [the reverse reaction, in which the stronger base, CN\(^{-}\) is acting as a base, DOES “go”]. Either way you look at it, the reactants are what are favored at equilibrium.)

7. (a) 2.9 x 10\(^{-9}\) M
(b) 1.9 x 10\(^{-33}\) M (as expected, the solubility is MUCH lower when one of the product ions [here, OH\(^{-}\)] is already present in solution from another source)

8. [OH\(^{-}\)] = 1.6 x 10\(^{-6}\) M

9. Q = (7.2/2)/(2.1/2)(1.3/2) = 5.27 > K (= 4.1). So since Q > K, there are “too many products” to be at equilibrium and net reaction will occur in the REVERSE direction (until equilibrium is established).

10. (a) BASIC because SCN\(^{-}\) is a weak base and K\(^{+}\) is neither an acid nor a base
(b) ACIDIC because NH\(_4\)\(^{+}\) is a weak acid and Cl\(^{-}\) is a negligible base (conjugate to the strong acid HCl)