1442 Answers to final review

1. \[11 \text{M} = 11 \text{ moles/L of solution} \quad \text{H}_2\text{SO}_4 = 11(98.1) = 1079.1 \text{ g/ L of solution}\]

1 L of solution has a mass of: \((1000 \text{ml})(1.59 \text{ g/ml}) = 1590 \text{g per Liter of solution}\)

Mass\% \text{H}_2\text{SO}_4 = \frac{\text{Part}}{\text{whole}} \times 100 = \frac{1079.1 \text{ g}}{1590 \text{g}} \times 100 = 67.9 \% \]

2. I. is least soluble non-polar, II most polar and smaller number of C, III middle because it is polar but more C then I. (Remember if a molecule contains three different elements it is going to be polar.)

3. \[P_{\text{solv}} = P_{\text{solv}}X_{\text{solv}} = (88.7 \text{ torr})(2.5 \text{ mol/ 2.5mol + .52mol}) = 73.9 \text{ torr}\]

4. First calculate the pressure of the vapor:
\[P_t = P^o_a X_a + P^o_b X_b\]
\[= (745\text{torr})(.425) + (290 \text{ torr})*(1-.425) = 483.4 \text{ torr}\]

Mole fraction of toluene above the solution = \(\frac{P_{\text{toluene}}}{P_{\text{total}}} = \frac{.575 \times 290}{483.4} = .345\)

5. Heat given off(exothermic) indicates a strong interaction between the solvent and the solute. this causes less of them to turn into gases and in effect causes the pressure to be less than expected.

6. \(\text{CaCl}_2\) breaks into three ions: one \(\text{Ca}^{2+}\) and two \(\text{Cl}^-\). The solution was \(.22\text{m X 3}\) that makes it \(.66\) in particles. The more the particles the more the change in freezing or boiling point.

7. \[\Delta T_b = \text{im}K_b = (1) \left[\left(\frac{40\text{g}}{160}\right) / .5 \text{Kg}\right] (2.8) = 1.4\]

\(\text{BP} = 81.0 + 1.4 = 82.4\) (boiling point elevation)

8. \(\Pi = \text{MRT}\)

\[\Pi = \left(\frac{\text{g/MM}}{\text{L}}\right)\text{RT} \quad (\text{substitue g/molar mass for moles})\]

\[\text{MM} = \text{gRT/L}\Pi = (1.73\text{g})(.0821 \text{ L atm/mol K})(298\text{K}) / (2.72 \text{ atm})(.1 \text{ L}) = 156\]

9. Gases in liquids increase solubility with high pressure and low temperature. Very different from solids.

10. \[\Delta T_b = \text{im}K_b = (3)\left[\left(\frac{30}{174.3}\right) / .1 \text{Kg}\right] (51) = 2.63\]
BP = 100 + 2.63 = 102.6

11. Heat of solution = heat of hydration + lattice energy
   -7.6 kJ = heat of hydration + 686 kJ
   heat of hydration = -694 kJ

12. See answer

13. \( \Pi = MRT \)
    \[ 7.8 = M \left( \frac{.0821}{\text{Latm/molK}} \right)(298K) \]
    M = .32M but it is .2 M so \( .32/.2 = 1.6 \). It needed to be 1.6 in order to make the pressure = 7.8 atm. MgSO_4 ionized into 2 particles, so i should be close to 2.

14. For the order of B use exp 1 and 2. Since A is constant it can be left out.
   \[ \frac{[0.015]}{[0.030]} = \frac{[0.010]}{[0.020]} \]
   \[ b = 1 \]

   For the order of A use exp 1 and 3
   \[ \frac{[0.033]}{[0.099]} = \frac{[0.010]}{[0.090]} \]
   \[ a = 2 \]

   Rate Equation = \( R = k[A]^2[B] \)

15. \( (0.082 \text{ M/s/1})(2 \text{ NH}_3 / 3 \text{ H}_2)^* = 0.055 \) * use coefficients in equation

16. Third order means the concentration is cubed. If you double the concentration: \( [2]^3 = \text{factor of 8} \)

17. \( t_{1/2} = \ln 2/k \)
   \[ 102s = \ln 2/k \]
   \[ k = .0068 \text{s}^{-1} \]

   \[ \ln[A]_0/[A]_t = kt \]
   \[ \ln (100/65) = .0068/s \text{ t} \]
   \[ t = 63.4 \text{ s} \]

18. \( 1/[A]_t \text{ vs t straight line = 2nd order, ln[A] vs t = first order, [A] vs t is zero order} \)
19. \[ \frac{1}{[A]} - \frac{1}{[A]_0} = kt \]
   \[ \frac{1}{[.08]} - 1[.1] = k(40\text{min}) \]
   \[ k = 0.0625/\text{min} \]

   \[ t_{1/2} = \frac{1}{k[A]_0} \]
   \[ t_{1/2} = \frac{1}{(0.0625\text{min}0[.1]} \]
   \[ t_{1/2} = 160 \text{ min} \]

20. 2 reactants = bimolecular, intermediate is formed in one step and consumes in another.

21. see answer

22. \[ \ln k_2/k_1 = -\frac{E_a}{R}(1/T_2 - 1/T_1) \]

   \[ \ln k_2/ 2.1 \times 10^5 \text{s}^{-1} = - \frac{135000 \text{J/mol}}{(8.314 \text{ J/molK})} [1/550 - 1/355] \]
   \[ k_2 = 2.3 \times 10^{12} \]

23. See answer

24. See answer. A catalyst starts as a reactant but cancels out. In other words it comes out just like it started.

25. rate = \( k[A]^2[B]^2 \)

   \[ k = \frac{[A]^2[B]^2}{\text{R}} = \frac{[\text{M}]^2[\text{M}]^2}{\text{M/L s}} = \text{answer d} \]

26. \( \text{Ni(s) + 4CO(g) \rightarrow Ni(CO)_4(g)} \) (Remember don't include solids)

   Eq
   \[ .01\text{M} \quad .005\text{M} \]

   \[ K = \frac{[\text{Ni(CO)}_4]}{[\text{CO}]^4} = \frac{[.005]/[.01]^4}{5 \times 10^5} \]

27. Since the reaction is exothermic, raising the temperature will shift it to the left or reactants. Since reactants are on the bottom of the equilibrium expression, that will make \( K \) get smaller.

28. \( K_p = K_c (RT) \Delta n_{\text{gas}} \)

   \[ 0.262 = K_c(0.0821 \times 1273)^{-1} \]
   \[ K_c = 27.4 \]

To get \( \Delta n \) subtract moles of gaseous products from moles of gaseous reactants using the coefficients for moles. In this case it is \( 1 - 2 = -1 \). \( C \) is not used since it is a solid.
29. When you reverse the equation it becomes the reciprocal: $1/K$. When you double the coefficients, it squares $K$: $1/K^2 = 2.4 \times 10^{-6}$

30. See answer. Remember to leave out solids and pure liquids.

31. $\text{R } 2H_2S \leftrightarrow 2H_2 + S_2$

<table>
<thead>
<tr>
<th></th>
<th>I</th>
<th>0.25M</th>
<th>0</th>
<th>0</th>
</tr>
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<tbody>
<tr>
<td>C</td>
<td>-2x</td>
<td>2x</td>
<td>x</td>
<td></td>
</tr>
<tr>
<td>E</td>
<td>0.25-2x</td>
<td>2x</td>
<td>x</td>
<td></td>
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</tbody>
</table>

$K = \left[ \frac{[H_2]^2[S_2]}{[H_2S]^2} \right]$ 

$9.3 \times 10^{-8} = \left[\frac{[2x]^2[x]}{[0.25 - 2x]^2}\right]$  (can ignore the -2x associated with 0.25) 

$9.3 \times 10^{-8} = 4x^3 / .0625$

\[x = 1.13 \times 10^{-3}\] Since H$_2$ is 2x then H$_2 = 2.3 \times 10^{-3}$

32. $Q = \left[ \frac{[NO]^2[Br_2]}{[NOBr]^2} \right] = \left[\frac{[.1]^2[.1]}{[.1]^2}\right] = .1$  \[K = 60.6\]

On a number line it looks this way: 0 .1 60.6. Starting at $Q$, you must move right to get to $K$ on the number line. That is also the direction the reaction will move. So the reaction will move to the right or forward. Remember, $Q$ and $K$ use the same expression. $Q$ is putting non-equilibrium amounts into the expression.

33. $\text{R } 2HI \leftrightarrow H_2 + I_2$

<table>
<thead>
<tr>
<th></th>
<th>I</th>
<th>.1</th>
<th>0</th>
<th>0</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td>-2X0.0135</td>
<td>+0.0135</td>
<td>+0.0135</td>
<td></td>
</tr>
<tr>
<td>E</td>
<td>.073</td>
<td>0.0135</td>
<td>0.0135</td>
<td></td>
</tr>
</tbody>
</table>

$K = \left[ \frac{[H_2][I_2]}{[HI]^2} \right] = \left[\frac{.0135^2}{.073^2}\right] = .0342$

34. See answer

35. See answer

36. $pOH = -\log[OH] = -\log[5.8 \times 10^{-3}] = 2.24$

$pH + pOH = 14$

\[pH = 11.76\]

37. $\text{R } HClO_4 + H_2O \rightarrow H_3O + ClO_4^-$

\[.035 \quad .035 \quad .035\] since strong acid it is 100% ionized
\[ \text{pH} = -\log[H^+] = - \log[0.035] = 1.46 \]

38. \( \text{Ba(OH)}_2 \rightarrow \text{Ba}^{2+} + 2 \text{OH}^- \)
\[
\begin{array}{ccc}
0.025 & 0.025 & 0.05 \\
\end{array}
\]
Since strong base it is 100% ionized

pOH = - log[0.05] = 1.30 the pH = 12.70

39. \( \text{HA} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{A}^- \)
\[
\begin{array}{ccc}
0.075 & 0 & 0 \\
-2.04 \times 10^{-4} & 2.04 \times 10^{-4} & 2.04 \times 10^{-4} \\
0.0748 & 2.04 \times 10^{-4} & 2.04 \times 10^{-4} \\
\end{array}
\]
\[ K = [\text{H}_3\text{O}^+][\text{A}^-] / [\text{HA}] = [2.04 \times 10^{-4}]^2 / 0.0748 = 5.6 \times 10^{-7} \]

40. \( \text{R} \text{C}_6\text{H}_5\text{COOH} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{C}_6\text{H}_5\text{COO}^- \)
\[
\begin{array}{ccc}
0.1 & 0 & 0 \\
-x & x & x \\
0.1-x & x & x \\
\end{array}
\]
\[ K = [\text{H}_3\text{O}^+][\text{C}_6\text{H}_5\text{COO}^-] / [\text{C}_6\text{H}_5\text{COOH}] \\
6.3 \times 10^{-5} = x^2 / 0.1 \quad (\text{ignore the x compared to 0.1 because K is small}) \]
x = 2.5 \times 10^{-3} = H^+ \quad \text{so pH} = 2.60

41. \( \text{R} \text{(CH}_3\text{)}_2\text{N} + \text{H}_2\text{O} \leftrightarrow \text{(CH}_3\text{)}_2\text{NH}^+ + \text{OH}^- \)
\[
\begin{array}{ccc}
0.15 & 0 & 0 \\
-x & x & x \\
0.15-x & x & x \\
\end{array}
\]
\[ K = [\text{H}_3\text{O}^+][\text{(CH}_3\text{)}_2\text{NH}^+] / [\text{CH}_3\text{N}] \\
6.3 \times 10^{-5} = x^2 / 0.15 \quad (\text{ignore the x associated with the 0.15 because K is small}) \\
x = 3.07 \times 10^{-3} = \text{OH}^- \quad \text{so pH} = 11.49
\]

42. \( \text{R} \text{H}_2\text{CO}_3 + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{HCO}_3^- \)
\[
\begin{array}{ccc}
0.1 & 0 & 0 \\
-x & x & x \\
0.1-x & x & x \\
\end{array}
\]
\[ K = [x^2] / [0.1] \quad (\text{ignore the x}) \\
4.5 \times 10^{-7} = [x^2] / [0.1] \quad x = 2.1 \times 10^{-4} \\
\]
\[ \text{R} \text{HCO}_3^- + \text{H}_2\text{O} \leftrightarrow \text{CO}_3^{2-} + \text{H}_3\text{O}^+ \]
\[
\begin{array}{ccc}
2.1 \times 10^{-4} & 0 & 2.1 \times 10^{-4} \\
-x & x & x \\
2.1 \times 10^{-4} -x & x & 2.1 \times 10^{-4} + x \\
\end{array}
\]
* the 2.1 X 10^{-4} values came from the first equilibrium problem

\[ K = \frac{[x][2.1 \times 10^{-4}]}{[2.1 \times 10^{-4}]} \] (ignore both x's)
\[ 4.7 \times 10^{-11} = x = [CO_3^{2-}] \]

\[ H_3O^+ = 2.1 \times 10^{-4} + 4.7 \times 10^{-11} = 2.1 \times 10^{-4} \] so pH = 3.67

43. Ka X Kb = 1 \times 10^{-14}
\[ Kb = 1 \times 10^{-14} / Ka = 1 \times 10^{-14} / 1 \times 10^{-10} = 1 \times 10^{-4} \]

44. \[ R \quad CN^- + H_2O \leftrightarrow HCN + OH^- \]
\[ I \quad .1 \quad 0 \quad 0 \]
\[ C \quad -x \quad x \quad x \]
\[ E \quad .1-x \quad x \quad x \]

\[ K_b = \frac{[x^2]}{[.1]} \] (ignore x)
*2.04 \times 10^{-5} = \frac{[x^2]}{[.1]}
\[ x = 1.43 \times 10^{-3} = OH^- \] so pH = 11.15

This is a Kb reaction since CN^- (the reactant) is a base. So you must change Ka Kb by using: \( Ka \times Kb = 1 \times 10^{-14} \).

45. With KCN since K^+ is form the strong base KOH and CN^- is from the weak acid HCN, CN^- will react with water. When you say strong, it means it will stay as an ion and will not react with water.
\[ CN^- + H_2O \leftrightarrow HCN + OH^- \] Since it ends with OH^- it is basic.

46. When decreasing the volume, the reaction moves to the side with the fewest moles of gases. Use the coefficients of the balanced equation to determine how many moles of gas on each side. In this example there are 3 moles of gas on the reactant side and 2 moles of gas on the product side. Remember to only consider gases.

47. See answer

48. \[ CH_3COOH + H_2O \leftrightarrow CH_3COO^- + H_3O^+ \]
\[ x \quad -5 \times 10^{-3} \quad 5 \times 10^{-3} \]
\[ 1.8 \times 10^{-5} = \frac{[5 \times 10^{-3}]^2}{[x-5 \times 10^{-3}]} \]
\[ x = 1.39 \]

If HCl is 5 \times 10^{-3} M the that is also the concentration of the H^+ and the Cl^- . So the problem has given the concentration of the H_3O^+ in the acetic acid solution. It is a fact that the H_3O^+ must equal the CH_3COO^- concentration. If you start with
x M acid, the $5 \times 10^{-3}$ must have changed to products. Thus the beginning concentration of acid is $x - 5 \times 10^{-3}$. The $5 \times 10^{-3}$ is usually the $x$ that we ignore, but since we know its value, we should not ignore it. But in reality, if we ignore the $-5 \times 10^{-3}$ and leave only $x$ on the bottom of the expression, it will give us the same answer.

49. The salt of a weak acid will always give you a basic solution.

50. HNO$_3$ strong acid, NH$_4$Cl salt that makes an acidic solution, KCl salt that makes a neutral solution, NaC$_2$H$_3$O$_2$ salt that makes a basic solution, KOH strong base

51. If you need help with writing net ionic equations go to the chemistry help topics on the website and look at Net Ionic Equations.

52. Remember, what must be left after mixing is a weak acid and its conjugate base. In II: NaC$_2$H$_3$O$_2$ and HC$_2$H$_3$O$_2$ and in IV the same.

53. R HC$_2$H$_3$O$_2$ + H$_2$O $\leftrightarrow$ C$_2$H$_3$O$_4^-$ + H$_3$O$^+$

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<tr>
<th></th>
<th>C$_2$H$_3$O$_4^-$</th>
<th>H$_3$O$^+$</th>
</tr>
</thead>
<tbody>
<tr>
<td>I</td>
<td>.107*</td>
<td>.0286</td>
</tr>
<tr>
<td>C</td>
<td>-x</td>
<td>x</td>
</tr>
<tr>
<td>E</td>
<td>.107-x</td>
<td>.0286 + x</td>
</tr>
</tbody>
</table>

$1.8 \times 10^{-5} = [0.0286][x] / 0.107$ (ignore both $x$'s)

$x = 6.73 \times 10^{-5} = H_3O^+$ so pH = 4.17

* Any time two volumes are mixed the M will change. You may use the formula $M_1V_1 = M_2V_2$ to get the new molarity.

54. R CN$^- + H^+$ $\rightarrow$ HCN

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<thead>
<tr>
<th></th>
<th>CN$^-$</th>
<th>HCN</th>
</tr>
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<tbody>
<tr>
<td>I</td>
<td>.208</td>
<td>.208</td>
</tr>
<tr>
<td>C</td>
<td>-.167</td>
<td>+.167</td>
</tr>
<tr>
<td>End</td>
<td>.041</td>
<td>.375</td>
</tr>
</tbody>
</table>

Now a buffer (weak acid and conjugate base) is left over. You may use the Henderson-Hasselbalch equation.

\[
pH = pK_a + \log \text{base/acid} \\
= - \log 4.9 \times 10^{-10} + \log .041 / .375 \\
= 8.35
\]

55. R H$^+$ + OH$^-$ $\rightarrow$ H$_2$O

<table>
<thead>
<tr>
<th></th>
<th>H$_2$O</th>
</tr>
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<tbody>
<tr>
<td>I</td>
<td>.0714</td>
</tr>
<tr>
<td>C</td>
<td>-.0571</td>
</tr>
<tr>
<td>End</td>
<td>.0143</td>
</tr>
</tbody>
</table>

End .0571 .0571 .0571
Since $H^+$ is left over $pH = -\log [.0143] = 1.85$

56. $\text{R} \quad \text{HF} + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{F}^-$
   I \quad 0.1 \quad 0.075 \quad 0
   C \quad -0.75 \quad -0.075 \quad +0.075
   E \quad 0.025 \quad 0 \quad 0.075

What is left is a buffer (weak acid and its conjugate base) so you may use the Henderson-Hasselbalch equation.

\[
pH = pK_a + \log \text{base/acid}
\]

\[
= -\log 3.5 \times 10^{-4} + \log \frac{0.075}{0.025} = 3.93
\]

57. The equivalence point is where stoichiometrically equal moles of both reactants have been added. In this problem we start with 50 ml of .3 M acetic acid. Since $M \times L = \text{moles}$, we start with .015 moles of acetic acid. Therefore we must add .015 moles of NaOH too. Since the NaOH is .2M : $L = \text{moles/M} = 0.015/.2 = .075L$ or 75 ml.

So we are mixing: 50 ml .3 M acetic acid + 75 ml of .2M NaOH

\[
\text{R} \quad \text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^- \rightarrow \text{H}_2\text{O} + \text{C}_2\text{H}_3\text{O}_2^-
\]
   I \quad 0.12 \quad 0.12 \quad 0
   C \quad -0.12 \quad -0.12 \quad +0.12
   End \quad 0 \quad 0 \quad 0.12

At the end of the titration acetate is left which is a base and needs a $K_b$.

\[
\text{R} \quad \text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \leftrightarrow \text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^- 
\]
   I \quad 0.12 \quad 0 \quad 0
   C \quad -x \quad x \quad x
   E \quad 0.12-x \quad x \quad x

\[
K_b = \frac{[x]^2}{.12} \; \text{(ignore the x)}
\]

\[
5.56 \times 10^{-10} = x^2/.12
\]

\[
x = 8.1 \times 10^{-4} = \text{OH}^- \; \text{so} \; pH = 8.91
\]

58. $\text{Cu}_3(\text{PO}_4)_{2(s)} \rightarrow 3 \text{Cu}^{2+} + 2 \text{PO}_4^{3-}$
   
   \[
   \frac{5.0 \times 10^{-8}}{3.3 \times 10^{-8}}
   \]

\[
K_{sp} = [\text{Cu}^{2+}]^3[\text{PO}_4^{3-}]^2 = [5.0 \times 10^{-8}]^3 [3.3 \times 10^{-8}]^2 = 1.4 \times 10^{-37}
\]

59. $\text{PbF}_2(s) \leftrightarrow \text{Pb}^{2+} + 2 \text{F}^-$
   
   \[
   x \quad x \quad 2x
   \]
Ksp = [Pb\(^{2+}\)] [F\(^{-}\)]^2
3.6 \times 10^{-8} = [x][2x]^2

x = 2.1 \times 10^{-3}

60. CaF\(_2\)(s) ⇔ Ca\(^{2+}\) + 2F\(^{-}\)

\[
x \quad x + .05* \\
2x
\]

Ksp = [.05][2x]^2 (ignore the x)

3.9 \times 10^{-11} = [.05][4x^2]

x = 1.4 \times 10^{-5} this is the molar solubility of CaF\(_2\).

* The additional .05 came from the .05 M Ca(NO\(_3\))\(_2\). When it ionized it produces one Ca\(^{2+}\) and 2 NO\(_3\)^{-}.

61. Ag\(_2\)SO\(_4\)(s) ⇔ 2 Ag\(^{+}\) + SO\(_4\)^{2-}

Ag\(^{+}\) concentration is .004 M (use the dilution equation)
SO\(_4\)^{2-} concentration is .0075

Q = [Ag\(^{+}\)]^2 [SO\(_4\)^{2-}] = [.04]^2[.0075] = 1.2 \times 10^{-5} Ksp = 1.2 \times 10^{-5}

Since Q = K the solution is at equilibrium(saturated) and there is no precipitate.

62. Lowering the pH is adding H\(^{+}\). In ( I ) the carbonate will be pulled out by the H\(^{+}\) causing the reaction to shift to the right. In (III) the F\(^{-}\) will be pulled out by the H\(^{+}\) causing the same effect.

63. See answer

64. Gas is more disordered than a liquid which is more disordered than a solid. If the moles of gas and liquids gets larger on the product side it is going to disorder or increased entropy.

65. See answer

66. Looking at the equation: \(\Delta G = \Delta H - T\Delta S\). If H is negative and S is negative (we know this because the reaction is going toward order) then the factor -T\(\Delta S\) is positive. We need that factor to be less than \(\Delta H\) so that \(\Delta G\) can be negative( this means the reaction is spontaneous). This will happen under low temperatures.
The lower $T$, the smaller $-T\Delta S$ will be.

67. $G$ is + because it is not spontaneous at $25^\circ C$. $H$ is - because heat must be lost for it to freeze, and $S$ is - because when it turns into a solid it will go toward order.

68. $\Delta G = \Delta H - T\Delta S$

\[ \Delta G = 29.24 \text{kJ} - T(0.0875 \text{kJ}) \]

$T = 334.2 \text{K}$ which is $61 \text{C}$

Boiling is an equilibrium process of $\Delta G$ is 0.

69. $\Delta S = \sum S_{\text{products}} - \sum S_{\text{reactants}}$

\[ \Delta S = [3 \times 210.6] - [220 + 240.5] = 171.3 \text{J} = 0.1713 \text{kJ} \]

\[ \Delta H = \sum H_{\text{products}} - \sum H_{\text{reactants}} \]

\[ \Delta H = [3 \times 90.4] - [81.6 + 33.8] = 155.8 \text{kJ} \]

\[ \Delta G = \Delta H - T\Delta S = [155.8] - (298)(0.1713) = 105 \text{kJ} \]

70. $\Delta G = \Delta G^0 + R T \ln Q$ ( $Q = [\text{NO}_2][\text{O}_2]/[\text{NO}][\text{O}_3]$ )

\[ \Delta G = -198 \text{kJ} + (0.008314 \text{kJ/mole K})(298) \ln[3 \times 10^{-6}][0.2]/[1 \times 10^{-6}][1 \times 10^{-7}] \]

\[ = -159 \]

71. $\Delta G = \sum G_{\text{products}} - \sum G_{\text{reactants}}$

\[ \Delta G = -32.9 - 209.2 \] ( remember $G$ for elements in standard state is 0 )

\[ = -242.1 \text{kJ} \]

\[ \Delta G = -RT\ln K \]

\[ -242.1 \text{kJ} = -(0.008314 \text{kJ})(298) \ln K \]

\[ K = 2.7 \times 10^{42} \]

72. III because Cu goes from 0 to +1 and S goes from 0 to -2. For a reaction to be a redox, one element goes up and one goes down.

73. An oxidizing agent must be reduced. Ni is not every going to be -1. In every other case there is an element with a high oxidation number allowing it to go down.

74. See answer. ( If you have trouble with this go to the website and open the Chemistry help topics link and look fore assigning oxidation numbers.)
75. P: in PH3 is it -3 and in PCl3 is it 3+.

76. S: In SO4\(^{2-}\) S is +6 and in SO2 it is +4.

77. The oxidizing agent is reduced. N goes from +5 to +1.

78. If you need help balancing redox equations go to the website and open Chemistry help topics and open the link balancing redox equations.

79. Same as previous question.

80. For a reaction to be spontaneous \( \Delta G \) must be negative, E (the voltage) must be positive, and K must be greater than one.

81. Look at the list of reduction potentials. The one with the most positive EMF (volts) is most likely to happen. You will notice that I\(^-\) and Br\(^-\) do not show up as reactants on the reduction potential list. In that case do not choose them.

82. Reverse the reduction potentials and you will get the oxidation potentials. Again, choose the one with the most positive EMF (volts). When you reverse the equations, the choices in the problem must be a reactant. In this case a, c, and d are not reactants when the reaction is reversed so do not choose from them.

83. Remember: Reduction occurs at the cathode. choice d is the only one in which an element (S in this case) goes down in oxidation number.

84. Ni is -.26 and Cd is -.4 so Ni is more likely to be reduced. That means it will be reduced or gain electrons. Therefore, the electrons must move toward the Ni.

85. Using the same process in 84 we find that the Zn is reduced. That means it goes from Zn\(^{2+}\) to Zn. Zn with no charge means a solid metal.

86. \( E_{\text{cell}} = E_{\text{oxidation(Ag)}} + E_{\text{reduction(Br)}} \)
   \( = -.8 + 1.09 = .29 \text{V} \) (remember when you get the oxidation potential you must change the sign of the volts.)

87. \( E_{\text{cell}} = E_{\text{oxidation(Co)}} + E_{\text{reduction(Ag)}} \)
   \( 1.08 = E_{\text{Co}} + .8 \) 
   \( E_{\text{oxidation(Co)}} = .28 \) so the reduction potential is -.28.

88. \( E_{\text{cell}} = \*E_{\text{cell}}^{o} -.0592V/n \log*Q \)
   \( = .23 -.0592V/2 \log \left[ \text{Fe}^{2+} \right]^{2} / \left[ \text{Fe}^{3+} \right]^{2} \left[ I^{-} \right]^{2} \)
   \( = .23 - .0592V/2 \log \left[ .1 \right]^{2} / \left[ 1.5 \right]^{2}[1.5]^{2} \)
*E^o_{cell} = E_{oxidation} + E_{reduction} = -.54 + .77 = .23

*Balanced Equation: $2 \text{Fe}^{3+} + 2I^- \rightarrow 2 \text{Fe}^{2+} + I_2$

89. $E_{cell} = E^0_{cell} - 0.0592 \frac{V}{n} \log \frac{[X]_{dil}}{[X]_{conc}}$

   = 0 - .0592V/2 \log [.1]/[.3] = .014V   The greater concentration cell is is reduced and is therefore the cathode.

90. $\Delta G = -nFE_{cell} = -*10(96500)(.15) = -144$

   * 10Cl^- to 5 Cl_2 is a change of 10 electrons
   * use table to get E oxidation + E reduction

91. $\log K = n*E^o_{cell} / 0.0592V$

   = (2)(.07)/.0592
   \[K = 2.3 \times 10^2\]

   * Use reduction potentials to get E oxidation + E reduction

92. $2\text{MnO}_4^- (\text{Mn is } +7) \rightarrow 2 \text{Mn}^{2+} (\text{Mn is } +2) 2 \text{ change by } 5 \text{ electrons} \text{ which is } 10$

   or $5 \text{H}_2\text{O}_2 (\text{O is } -1) \rightarrow 5 \text{O}_2 (\text{O is zero}) \text{ is a change of } 10$

   You may either use the oxidation half-reaction or the reduction half-reaction to determine the change in the electrons.

93. $i = nFe$

   $i = (40g/158.5g/mol)(96500 \text{C})(3\text{electrons}) = 7.31 \times 10^4 \text{C}$

   $i = \text{current} \quad t = \text{time seconds} \quad i = \text{coulombs}$

94-110 read notes and see answers on sheet.

As an example in 95 CH_3OH = hydrogen bonding since H is attached to O

NH_3 = Hydrogen bonding since H is attached to N

H_2S = dipole-dipole since it is a polar molecule

Kr = London Dispersion since it is non-polar