1) What is required to make an effective buffer solution?

A WEAK acid and a salt which contains the CONJUGATE BASE of the weak acid.

2) What effect does a common ion have on the solubility of a salt?

The common ion will REDUCE the solubility of the salt by shifting the equilibrium position to the left.

3) What is the net ionic equation for a strong acid / strong base titration?

$H^+ + OH^- \rightarrow H2O$ (ALWAYS!!!)

4) Know the basic rules for assigning oxidation numbers.

Read through pages 134 and 135!!!!!!!

5) What is the relationship between cell potential, work, and spontaneity?

Only SPONTANEOUS reactions can perform work, only cells with a POSITIVE potential are spontaneous, ergo, they can do work.

6) A buffer solution is prepared from 0.75 M HA ($Ka = 3.2 \times 10^{-7}$) and 0.15 M NaA. What is the pH of the solution?

pH = pKa + log[base]/[acid]

$pH = -\log(3.2 \times 10^{-7}) + \log (0.15/0.75) = 5.80$

7) How do you determine which indicator to use in a titration?

The indicator should change color in the pH range of the equivalence point.

8) Be able to identify the distinguishing features of any titration curve.

Starting point pH (range) for SA/SB, WA/SB, WB/SA

Equivalence point pH (range) for SA/SB, WA/SB, WB/SA

9) What is the solubility product expression for Na_2CO_3?

$Ksp = [Na^+]^2[CO_3^{2-}]$

10) What is the K_{sp} expression (in terms of molar solubility) for Na_2CO_3?

$Ksp = (2s)^2(s) = 4s^3$

11) The molar solubility (s) of Na_2CO_3 is 4.7×10^{-5} M. Calculate the Ksp value.

$Ksp = (2s)^2(s) = 4s^3 = (4)(4.7 \times 10^{-5})^3 = 4.15 \times 10^{-13}$

12) Explain how the pH of a solution can affect the solubility of a salt.

In general, the H^+ (acidic solution) or OH^- (basic solution) will act like a common ion and shift the equilibrium position and therefore the amount of salt which will dissolve. You have to know what ions the salt produces in solution before you can determine the overall affect.

13) What is the oxidation number of Sulfur (S) in HSO_4^-?

(1)(+1) + (1)(X) + (4)(-2) = -1 X = +6
H S O

14) In an electrochemical cell, how do you determine which side is the anode and which is the cathode?

The ANODE side is the OXIDATION reaction (loses electrons!!)

The CATHODE side is the RECUCTION reaction (gains electrons!!)

15) What is the proper format for writing the cell notation?

Anode || Cathode

Anode is always on the LEFT!!!!!

16) Determine the cell potential for the following reaction.

$$Pt^{2+}\ (aq) + Cr\ (s) \rightarrow Cr^{3+}\ (aq) + Pt\ (s)$$

(ANODE) $Cr^{3+}\ (aq) + 3\ e^- \rightarrow Cr\ (s)$ $E°_{red} = -0.74$ V
(CATHODE) $Pt^{2+}\ (aq) + 2\ e^- \rightarrow Pt\ (s)$ $E°_{red} = +1.20$ V

$E°_{cell} = E°_{cathode} - E°_{anode} = (1.20\ V) - (-0.74\ V) = 1.94$ V

17) Calculate the cell potential (E_{cell}) made from 0.72 M Co^{2+} and 0.07 M Fe^{2+} at 298 K.

$$Co\ (s) + Fe^{2+}\ (aq) \rightarrow Co^{2+}\ (aq) + Fe\ (s) \qquad E°_{cell} = +2.63\ V$$

$$E_{cell} = E°_{cell} - \frac{0.0257V}{n}\ln Q$$ n= number of electrons

$$E_{cell} = (2.63V) - \frac{0.0257V}{2}\ln\frac{0.72}{0.07} = 2.60V$$

18) How do you use Q and Ksp to determine precipitate formation?

See Example 16.10 in book (pg 724) and example from notes!!!!!

19) Know how to balance electrochemical reactions in both acidic and basic solutions.

i. Write out net ionic equation
ii. Write out each half reaction
iii. Balance each half reaction (Mass Balance (can add H_2O and H^+) BEFORE Charge Balance)
iv. Sum up the half reactions (equalize number of electrons)
v. If in a BASIC solution, add OH^- to form water with H^+, cancel waters
vi. Recheck Mass and Charge balance (may need to add OH^- in last step for BASIC)

20) Explain the parts of a galvanic cell and what each part does.

See diagram from class and know the following:

Anode side
Cathode side
Salt bridge
Direction of electron flow