Answers Test 4 Chapters 20 and 21 Silberberg

1. c

2. (1g CH_4/1)(1 mol CH_4/ 16g)(891 kJ/ 1 mole CH_4)* = 55.7 kJ

   The conversion factor 891 kJ/ 1 mole comes from the coefficient in the balanced equation. You will notice that the answer is usualy expressed as a positive number. If you are given the choice of -55.7 kJ pick that.

3. c

4. 2, 3

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

   $= -403.13\text{kJ/mol} - (273\text{ K})(-.1893\text{ kJ/mol.K}) = -351.45\text{ kJ/mole}$

   Notice that S will always be in J so you will need to change it to kJ.

6. $\Delta G = -RT \ln K$

   $\ln K = \Delta G / -RT = 217\text{ kJ/mole} / -(.008314\text{ kJ/mol.K})(298\text{ K})$

   $\ln K = -87.586$

   $K = 9.16 \times 10^{-39}$

   Notice that R is in J and had to be converted to kJ. To find K, use 2nd then ln on the calculator.

7. $\Delta S = \sum S_{products} - \sum S_{reactants} = [2NO_2 + 3\ H_2O] - [2NH_3 + 7/2\ O_2]$

   $= [2(240.40 + 3(188.7)] - [\ (2)192.5 + (7/2)\ 205]$

   $= [1046.1] - [1102.5] = -56.4\text{ J}$

8. Assign oxidation numbers first:

   $XeF_2 + BrO_3^- \rightarrow Xe + 2\ HF + BrO_4^-$

   2+ 1- +5 2- 0 1+ 1- 7+ 2-

   Xe is the oxidizing agent and Br is oxidized

9. $MnO_4^- + 8H^+ + 5e \rightarrow Mn^{2+} + 4\ H_2O$

   $5(\ Fe^{2+} \rightarrow Fe^{3+} + e\ )$

   $MnO_4^- + 8\ H^+ + 5\ Fe^{2+} \rightarrow Mn^{2+} + 4\ H_2O + 5\ Fe^{3+}$ Balanced in acid

In Base: $8\ OH^-$ → $8\ OH^-$

$$MnO_4^- + 8\ H_2O + 5\ Fe^{2+} \rightarrow Mn^{2+} + 4\ H_2O + 5\ Fe^{3+} + 8\ OH^-$$

Combine water $MnO_4^- + 4\ H_2O + 5\ Fe^{2+} \rightarrow Mn^{2+} + 5\ Fe^{3+} + 8\ OH^-$

10. C

11. A

12. $E_{cell} = E_{oxidation} + E_{reduction}$
$= 1.63 + -1.21 = 0.42$ V

Both of the equations given are reduction potentials. The more positive one (CdS) will remain reduced and the other will have to be reversed and oxidized. When oxidized, the sign is changes so the oxidation potential for Ti is 1.63.

13. $E_{cell} = E_{oxidation} + E_{reduction}$
$= -.8 + .99 = .19$ V if the E is positive, it is spontaneous.

14. $E = E^o - .0592V/n \log Q$

$= 0.3V - .0592\ V / (2) \log [Fe^{2+}] / [Sn^{2+}]$
$= 0.3V - .0592\ V / (2) \log [.5] / [.001]$

$= 0.3\ V - .08 = .22\ V$

To get the expression for Q you must reverse the Fe^{2+} reaction:
$Sn^{2+} + Fe \leftrightarrow Fe^{2+} + Sn$

To get 0.3V for E use $E_{cell} = E_{oxidation} + E_{reduction}$

15. $\Delta G = -nFE = -(1)(96,500C)(0.5\ V) = -48250$ J or - 4.8 kJ [1 V = 1 J/C]

n stands for the moles of electrons transferred in the balance equation. In this eqution Mn^{3+} is both oxidized and reduced:

$Mn^{3+} \rightarrow Mn^{2+}$ change of 1 electron
$Mn^{3+} \rightarrow MnO_2$ change of 1 electron (Mn in MnO_2 is 4+)

To determine n you can use either the oxidation half-reaction or the reduction half-reaction.

16. $it = nFe$

(10 amps)(600 s) = n (96,500C)(2)

n = .062 moles of Ni

(.062 mol Ni/1)(58.7g/mol Ni) = 3.6 g Ni

17. $\log K = nE^o / 0.0592V = (2)(0.01\ V) / 0.0592\ V$

K = 2.17

n is 2 because: $H_2 \rightarrow 2H^+$ or $NO_3^- \rightarrow NO_2^-$ is a change of 2 electrons.

18. b

19. b

20. a

21. c

22. e

23. e

24. c

$it = nFe$
(3.86 amps)(972s) = 1.52/MM(96500C)(1)
MM = 39

In the equation substitute n, moles, with g/molar mass. e is 1 because all of the elements are in column I on the periodic table and it takes one electron to bring them to the zero oxidation state which is the elemental form.