

Answers Test 4 Chapters 20 and 21 Silberberg

1. c

2. $(1\text{g CH}_4/16\text{g})(1\text{ mol CH}_4/16\text{g})(891\text{ kJ/ 1 mole CH}_4)^* = 55.7\text{ kJ}$

The conversion factor 891 kJ/ 1 mole comes from the coefficient in the balanced equation. You will notice that the answer is usually 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. $\Delta G = -RT \ln K$

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

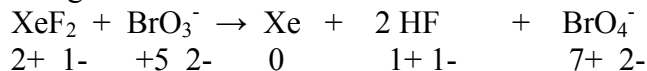
$$\ln K = -87.586$$

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

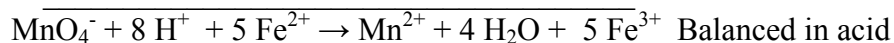
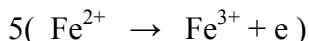
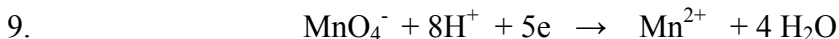
Notice that R is in J and had to be converted to kJ. To find K, use 2nd then ln on the calculator.

7. $\Delta S = \sum S_{\text{products}} - \sum S_{\text{reactants}} = [2\text{NO}_2 + 3\text{ H}_2\text{O}] - [2\text{NH}_3 + 7/2\text{ O}_2]$
 $= [2(240.40 + 3(188.7))] - [(2)192.5 + (7/2) 205]$
 $= [1046.1] - [1102.5] = -56.4\text{ J}$

8. Assign oxidation numbers first:



Xe is the oxidizing agent and Br is oxidized



In Base: $8 \text{ OH}^- \rightarrow 8 \text{ OH}^-$



Combine water $\text{MnO}_4^- + 4 \text{ H}_2\text{O} + 5 \text{ Fe}^{2+} \rightarrow \text{Mn}^{2+} + 5 \text{ Fe}^{3+} + 8 \text{ OH}^-$

10. C

11. A

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

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

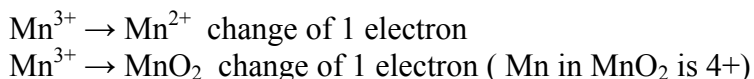
14. $E = E^{\circ} - .0592\text{V}/n \log Q$
 $= 0.3\text{V} - .0592 \text{ V} / (2) \log [\text{Fe}^{2+}] / [\text{Sn}^{2+}]$
 $= 0.3\text{V} - .0592 \text{ V} / (2) \log [.5] / [.001]$
 $= 0.3 \text{ V} - .08 = .22 \text{ V}$

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

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

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

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



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

16. $it = nFe$

$$(10 \text{ amps})(600 \text{ s}) = n (96,500\text{C})(2)$$

$$n = .062 \text{ moles of Ni}$$

$$(.062 \text{ mol Ni}/1)(58.7\text{g/mol Ni}) = 3.6 \text{ g Ni}$$

17. $\log K = nE^\circ / 0.0592\text{V} = (2)(0.01 \text{ V}) / 0.0592 \text{ V}$

$$K = 2.17$$

n is 2 because: $\text{H}_2 \rightarrow 2\text{H}^+$ or $\text{NO}_3^- \rightarrow \text{NO}_2^-$ is a change of 2 electrons.

18. b

19. b

20. a

21. c

22. e

23. e

24. c

$$it = nFe$$

$$(3.86 \text{ amps})(972\text{s}) = 1.52/\text{MM}(96500\text{C})(1)$$

$$\text{MM} = 39$$

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.