AP CHEMISTRY REVIEW
(problems from Zumdahl 5th edition)

Ch. 4: Reactions

p180

1. (11.) Show how each of the following strong electrolytes "breaks up" into its component ions upon dissolving in water.
   a. $MgCl_2$
   d. $(NH_4)_2SO_4$

2. (17) .d) Calculate the concentration of all ions present in the solution of the strong electrolyte $2.0x10^{-3}$ M $Al_2(SO_4)_3$.

3. (#29) When the following solutions are mixed together, what precipitate (if any) will form?
   a. $BaCl_2$*(aq)* + Na_2SO_4*(aq)*
   d. $NaOH$*(aq)* + $Fe(NO_3)_3$*(aq)*

p182

4. (#43.) A 100.0mL aliquot of 0.200M aqueous potassium hydroxide is mixed with 100.0mL of 0.200M aqueous magnesium nitrate.
   a. Write a balanced chemical equation for any reaction that occurs.
   b. What precipitate forms?
   c. What mass of precipitate is produced?
   d. Calculate the concentration of each ion remaining in solution after precipitation is complete.

5. (45) a) Write the balanced molecular, complete ionic, and net ionic equations for this acid base reaction:

   $HClO_4$*(aq)* + $Mg(OH)_2$*(aq)* →

6. (57) j) Assign Oxidation states for all atoms in this compound:

   $Na_2C_2O_4$

7. (63) c) Balance the following oxidation-reduction reaction that occurs in acidic solution:

   As_2O_3*(s)* + NO_3^- *(aq)* → H_3AsO_4*(aq)* + NO*(g)*

8. (65) c) Balance the following oxidation-reduction reactions that occurs in basic solution:

   NO_2^- *(aq)* + Al*(s)* → NH_3*(g)* + AlO_2*(aq)*

Ch. 7: Atomic Structure

p340

9. (# 31.) Why do the successive ionization energies of an atom always increase?

10. (35) Does the minimization of electron-electron repulsions correlate with Hund's rule and why?

11. (55) It takes 7.21 x 10^{-19}J of energy to remove an electron from an iron atom. What is the maximum wavelength of light that can do this?

P341

12. (# 59.) What are the possible values for the quantum numbers n, l, and m_l?

13. (61) Which of the following sets of quantum numbers are not allowed in the hydrogen atom? For the sets of quantum numbers that are incorrect, state what is wrong in each set.
    a. $n = 2, l = 1, m_l = -1$
    b. $n = 2, l = 1, m_l = -1$
    c. $n = 2, l = 1, m_l = -1$
    d. $n = 2, l = 1, m_l = -1$

14. (87) In each of the following sets, which atom or ion has the smallest radius?
    a. Li, Na, K
    b. P, As
    c. O+, O, O-

15. (93) The first ionization energies of As and Se are 0.947 and 0.941 MJ/mol, respectively. Rationalize these values in terms of electron configurations.

Ch 16: Thermochemistry

p829

16. (# 25.) Predict the sign ΔS_{surr} for the following processes.
    a. $H_20(l) \rightarrow H_20(g)$
    b. $CO_2(g) \rightarrow CO_2$ (s)

17. (31) For ammonia (NH_3), the enthalpy of fusion is 5.65kJ/mol and the entropy of fusion is 28.9 J/K• mol.
    a. Will $NH_3(s)$ spontaneously melt at 200 K?
    b. What is the approximate melting point of ammonia.

18. (39) For the reaction

$$C2H_2(g) + 4F_2\ (g) \rightarrow 2CF_4(g) + H_2\ (g)$$

ΔS° is equal to –358 J/K. Use this value and data from Appendix 4 to calculate the value of S° for $CF_4(g)$.

19. (49) Given the following data:

$S(s) + 3/2\ O_2(g) \rightarrow SO_3(g)$ $\quad \Delta G° = -371$ kJ

$2SO_2(g) + O2(g) \rightarrow 2SO_3$ $\quad \Delta G° = -142$ kJ

Calculate Δ G° for the reaction:

$6C(s) + 3H_2(g) \rightarrow C_6H_6(l)$

Ch. 5: Gases

p233

20. (# 35.) The steel reaction vessel of a bomb calorimeter, which has a volume of 75.0 mL, is charged with oxygen gas to a pressure of 145 atm at 22° C. Calculate the moles of oxygen in the reaction vessel.

p234

21. (# 55.) Hydrogen cyanide is prepared commercially by the reaction of methane, CH4(g), ammonia, NH3(g), and oxygen, O2(g), at high temperature. The other product is gaseous water.
    a. Write the chemical equation for the reaction.
    b. What volume of HCN(g) can be obtained from 20.0 L $CH_4(g)$, 20.0 L $NH_3(g)$, and 20.0 L $O_2(g)$? The volumes of all gases are measure at the same temperature and pressure.

22. (65) The partial pressure of CH4(g) is .175 atm and that of O2(g) is 0.250 atm in a mixture of the gases.
    a. What is the mole fraction of each gas in the mixture?
    b. If the mixture occupies a volume of 10.5 L at 65°C, calculate the total number of moles of gas in the mixture.
    c. Calculate the number of grams of each gas in the mixture.

p235

23. (# 101.) A 2.00-L sample of $O_2(g)$ was collected over water at a total pressure of 785 torr and 25°C. When the $O_2(g)$ was dried (water vapor removed), the gas

had a volume of 1.94 L at 25°C and 785 torr. Calculate the vapor pressure of water at 25°C.

Ch. 8&9: Bonding and Molecular Geometry

P406

24. (#35.) For each of the following groups, place the atoms and/or ions in order of decreasing size.
    a. Cu, Cu^{+}, Cu^{2+}
    b. Ni^{2+}, Pd^{2+}, Pt^{2+}

P407

25. (#39.) Which compound in each of the following pairs of ionic substances has the most exothermic lattice energy? Justify your answers.
    a. NaCl, KCl
    e. NaCl, Na_2O

26. (#47.) Use bond energy values (Table 8.4) to estimate ΔH for each of the following reactions in the gas phase.
    a. $H_2 + Cl_2 \rightarrow 2HCl$
    b. $N{\equiv}N + 3H_2 \rightarrow 2NH_3$
    c.

P408

27. (#55.) Consider the following reaction:

$$H_2(g) + O_2(g) \rightarrow H_2O_2(g) \qquad \Delta H = -153 \text{ kJ}$$

Given that the H_2 bond energy is 432 kJ/mol, the O_2 bond energy is 495 kJ/mol, and the OH bond energy is 467 kJ/mol, estimate the bond energy for the oxygen-oxygen single bond.

28. (#63.) Write Lewis structures that obey the octet rule for each of the following.
    a. HCN
    b. PH_3
    c. $CHCl_3$
    d. NH_4^{+}

P409

29. (#79.) Predict the molecular structure and bond angles for each molecule or ion in Exercise 63.

P410

30. (#85.) Predict the molecular structure (including bond angles) for each of the following.
    a. $XeCl_2$
    b. ICl_3
    c. TeF_4
    d. PCl_5

P444

31. (#23.) For each of the following molecules, write the Lewis structure(s), predict the molecular structure (including bond angles), and give the expected hybrid orbitals on the central atom, and predict he overall polarity.
    b. NF_3
    i. KrF_4
    j. SeF_6
    l. IF_3

Ch. 10: Liquids and Solids

P504

32. (#35.) Identify the most important types of interparticle forces present in the solids of each of the following substances.
    a. Ar
    b. HCl
    c. HF
    d. $CaCl_2$
    e. CH_4
    f. CO
    g. $NaNO_3$

P508

33. (#81.) In Breckenridge, Colorado, the typical atmospheric pressure is 520 torr. What is the boiling point of water (ΔH_{vap} = 40.7 kJ/mol) in Breckenridge?

34. (#87.) How much energy does it take to convert 0.500 kg ice at -20°C to steam at 250°C? specific heat capacities: ice, 2.1 J/g°C; liquid, 4.2 J/g°C; steam, 2.0 J/g°C, ΔH_{vap} = 40.7 kJ/mol, ΔH_{fus} = 6.02 kJ/mol.

Ch. 11: Solutions

P552

35. (#9.) A solution is prepared by dissolving 125 g sucrose ($C_{12}H_{22}O_{11}$) in enough water to produce 1.00 L solution. What is the molarity of this solution?

36. (#13.) What volume of 0.25 *M* HCl solution must be diluted to prepare 1.00 L of 0.040 *M* HCl?

P553

37. (#27.) A solution is prepared by dissoling 50.0 g cesium chloride (CsCl) in 50.0 g water. The volume of the solution is 63.3 mL. Calculate the mass percent, molarity, molality, and mole fraction of the cesium chloride.

P554

38. (#41.) What factors cause one solute to be more strongly attracted to water than another? For each of the following pairs, predict which substance would be more soluble in water.
    a. CH_3CH_2OH or $CH_3CH_2CH_3$
    b. $CHCl_3$ or CCL_4
    c. CH_3CH_2OH or $CH_3(CH_2)_{14}CH_2OH$

39. (#47.) A solution is prepare by mixing 50.0 g glucose ($C_6H_{12}O_6$) with 600.0 g water. What is he vapor pressure of this solution at 25°C? (At 25°C the vapor pressure of pure water is 23.8 torr. Glucose is a nonelectrolyte.)

p555

40. (#59.) A solution is prepared by dissolving sucrose ($C_{12}H_{22}O_{11}$) in 175 g water. Calculate the boiling points of this solution. Sucrose is a nonelectrolyte.

41. (#61.) What mass of glycerin ($C_3H_8O_3$), a nonelectrolyte, must be dissolved in 200.0 g water to give a solution with a freezing point of –1.50°C?

42. (#67.) a. Calculate the freezing-point depression and osmotic pressure at 25°C of an aqueous solution containing 1.0 g/L of a protein (molar mass = 9.0×10^4 g/mol) if the density of the solution is 1.0 g/cm^3.
    b. Considering your answer to part a, which colligative property, freezing-point depression or osmotic pressure, would be better use to determine the molar masses of large molecules? Explain.

P556

43. (#71.) Consider the following solutions:
    0.010 *m* Na_3PO_4 in water
    0.020 *m* $CaBr_2$ in water
    0.020 *m* KCl in water
    0.020 *m* HF in water (HF is a weak acid.)

    a. Assuming complete dissociation of the soluble salts, which solution(s) would have the same boiling point a 0.040 *m* $C_6H_{12}O_6$ in water? $C_6H_{12}O_6$ is an electrolyte.
    b. Which solution would have the highest vapor pressure at 28°C?
    c. Which solution would have the largest freezing-point depression?

Ch. 12: Kinetics

P. 602

44. (#11) Why is the rate of a reaction affected by each of the following?
    a. frequency of collisions
    b. kinetic energy of collisions

c. orientation of collisions

45. (#23) The reaction was studied at -10°C. The following results were obtained where Rate = $-\Delta[Cl_2]/\Delta t$

$$2NO(g) + Cl_2(g) \rightarrow 2NOCl(g)$$

$[NO]_0$ *(mol/L)*	$[Cl_2]_0$ *(mol/L)*	Initial Rate *(mol/L • min)*
0.10	0.10	0.18
0.10	0.20	0.36
0.20	0.20	1.45

a. What is the rate law?
b. What is the value of the rate constant?

P. 604

46. (#29) The decomposition of hydrogen peroxide was studied, and following data were obtained as a particular temperature:

Time (s)	*$[H_2O_2]$ (mol/L)*
0	1.00
120 +--1	0.91
300 +--1	0.78
600 +--1	0.59
1200 +--1	0.37
1800 +--1	0.22
2400 +--1	0.13
3000 +--1	0.082
3600_+--1	0.050

Assuming that

Rate = $-\Delta[H_2O_2]/\Delta t$

determine the rate law, the integrated rate law, and the value of the rate constant. Calculate the *$[H_2O_2]$* at 4000. *s* after the start of the reaction.

P. 605

47. (#45) Write the rate laws for the following elementary reactions.

a. $CH_3NC(g) \rightarrow CH_3CN(g)$
b. $O_3(g) + NO(g) \rightarrow O_2(g) + NO_2(g)$

P. 606

48. (#59) One mechanism for the destruction of ozone in the upper atmosphere is

$$O_3(g) + NO(g) \rightarrow NO_2(g) + O_2(g)$$
$$NO_2(g) + O(g) \rightarrow NO(g) + O_2(g)$$
Overall reaction: $O_3(g) + O(g) \rightarrow 2\ O_2(g)$

a. Which species is a catalyst?
b. Which species is an intermediate?
c. E_a for the uncatalyzed reaction?

Ch. 13: Chemical Equilibrium

P. 650

49. (#19b) Write the equilibrium expression (K) for each of the following gas-phase reactions.

b. $N_2O_4(g) \leftrightarrow 2NO_2\ (g)$

50. (#21.) Calculate the values of K for the following reactions at this temperature.

a. $1/2N_2\ (g) + 3/2\ N_2\ (g) \leftrightarrow NH_3\ (g)$

b. $2NH_3\ (g) \leftrightarrow N_2\ (g) + 3H_2\ (g)$

p. 651

51. (#29.) At 25 degrees C, K=3.7 x 10^9 for the reaction $CO\ (g) + Cl_2(g) \leftrightarrow COCl_2\ (g)$

52. (#31.) Write the expression for K for the following reaction: $NH_4NO_3(s) \leftrightarrow N_2O(g) + 2H_2O(g)$

53. (#37.) The equilibrium constant, K = 2.4 x 10^3 at a certain temperature for the reaction

$$2NO(g) \leftrightarrow N_2(g) + O_2(g)$$

For which of the following sets of conditions is the system at equilibrium? For those that are not at equilibrium, in which direction will the system shift?

a. A 1.0L Flask contains 0.024 mol NO, 2.0 mol N_2, and 2.6 mol O_2.

b. A 2.0L flask contains 0.032 mol NO, 0.62 mol N_2, and 4.0 mol O_2.

c. A 3.0L flask contains 0.060 mol NO, 2.4 mol N_2, and 1.7 mol O_2.

p. 652

54. (#45.) At a particular temperature, 12.0 mol of SO_3 is placed into a 3.0L rigid container, and the SO_3 dissociates by the reaction

$$2SO_3(g) \leftrightarrow 2SO_2(g) + O_2(g)$$

At equilibrium, 3.0 mol of SO_2 is present. Calculate K for this reaction.

55. (#57.) A sample of solid ammonium chloride was placed in an evacuated container and then heated so that it decomposed to ammonia gas and hydrogen chloride gas. After heating, the partial pressure of NH_3 in the container was found to be 2.2 atm. Calculate K_p at this temperature for the decomposition reaction

$$NH_4Cl(s) \leftrightarrow NH_3(g) + HCl(g)$$

56. (#63.) In which direction will the position of the equilibrium be shifted for each of the following changes?

$$2HI(g) \leftrightarrow H_2(g) + I_2(g)$$

a. $H_{2(g)}$ is added.

b. $I_{2(g)}$ is removed.

c. $HI_{(g)}$ is removed.

d. Some $Ar_{(g)}$ is added.

e. The volume of the container is doubled.

The temperature is decreased (the reaction is exothermic).

Ch. 14: Acids and Bases

P. 712

57. (#25.) Consider the reaction of acetic acid in water.

$$CH_3CO_2H\ (aq) + H_2O\ (l) \leftrightarrows CH_3CO_2^-(aq) + H_3O^+(aq)$$

Where $K_a = 1.8 \times 10^{-5}$.

a. Which two bases are competing for the proton?
b. Which is the stronger base?
c. In light of your answer to (b) why do we classify the acitate ion ($CH_3CO_2^-$) as a weak base? Use an appropriate reaction to justify your answer.

58. (#33.) Use table 14.2 to order the following from strongest to weakest acid.
H_2O (1×10^{-14}), HNO_3 (ka>1), HOCl (3.5×10^{-8}), NH_4^+(5.6×10^{-10})

P. 713

59. (#63.) Calculate the pH of a solution containing a mixture of 0.10 *M* HCl and 0.10 *M* HOCL.

60. (#65.) Calculate the percent dissociation of the acid in the following solution.
b. 0.050 *M* acetic acid.

61. (#87.) What are the major species present in a 0.150 *M* NH_3 solution. Calculate the $[OH^-]$ and the pH of the solution.

62. (#89.) Calculate [OH-], [H+], and the pH of 0.20M solution of the following amine:

a. Triethylamine $[(C_2H_5)3N$, Kb= 4.0×10^{-4}

Ch. 15: Solution Equilibrium

p. 783

63. (#53) How many moles of NaOH must be added to 1 L of 2 M $HC_2H_3O_2$ to produce a solution buffered at each pH.

a) pH = pK_a

b) pH = 4.00

c) pH = 5.00

p. 784

64. (# 65.) Calculate the pH at the halfway point and at the equivalence point of each of the following of the following titrations.

a) 100.0 ml of 0.10 M HC7H5O2 ($K_a = 6.4 \times 10^{-5)}$ titrated

b) 100.0 ml of 0.10 *M* $C_2H_5NH_2$ ($K_b = 5.6 \times 10^{-4}$) titrated by 0.20 *M* HNO_3

65. (69) Two drops of indicator HIn ($K_a = 10 \times 10^{-9}$), where HIn is yellow and In- is blue, are placed in 100.0 mL of 0.10 M HCl.

a. What color is the solution initially?

b. The solution is titrated with 0.10 M NaOH. At what pHs will the color change (yellow to greenish yellow) occur?

c. What colors will the solution be after 200.0 mL of NaOH has been added?

P. 785

66. (93) Calculate the solubility of solid $Ca_3(PO_4)_2$ ($K_{sp} = 1.3 \times 10^{-32}$) in a 0.20 M Na_3PO_4 solutions.

67. (97) Will a precipitate form when 75 mL of .020 M $BaCl_2$ and 125 mL of .040 M Na_2SO_4 are mixed together.

K_{SP} problem from AP Chemistry Test 2001

68. (1) Answer the following questions relating to the solubility of the chlorides of silver and lead.

A. At 10°C, 8.9×10^{-5} g of AgCl(s) will dissolve in 100 mL of water.

(i) Write the equation for the dissociation of AgCl(s) in water.

(ii) Calculate the solubility, in mol L^{-1}, of AgCl(s) in water at 10°C.

(iii) Calculate the value of the solubility product constant, K_{sp} for the AgCl(s) at 10°C.

B. At 25°C, the value of K_{sp} for $PbCl_2(s)$ is 1.6×10^{-5} and the value of K_{sp} for AgCl(s) is 1.8×10^{-10}

(i) If 60.0 mL of 0.0400 M NaCl(aq) is added to 60mL of 0.0300 M $Pb(NO_3)_2(aq)$, will a precipitate form? Assume that volumes are additive. Show calculations to support your answer.

(ii) Calculate the equilibrium value of [Pb2+(aq)] in 1.0L of saturated PbCl2 solution to which .250 moles of NaCl(s) has been added. Assume that no volume change occurs.

(iii) If 0.100 M NaCl(aq) is added slowly to a beaker containing both 0.120 M AgNO3(aq) and 0.150 M $Pb(NO_3)_2$ (aq) at 25°C, which will precipitate first, AgCl(s) or $PbCl_2(s)$? Show calculations to support your answer.

$$2\ H_2S(g) \Delta\ 2\ H_2S(g) + S_2(g)$$

69. When heated, hydrogen sulfide gas decomposes according to the equation above. A 3.40 g sample of $H_2S(g)$ is introduced into an evacuated rigid 1.25 L container. The sealed container is heated to 483 K, and 372 x 10^{-2} mol of $S_2(g)$ is present at equilibrium.

    a) Write the expression for the equilibrium constant, K_c, for the decomposition reaction represented above.

    b) Calculate the equilibrium concentration, in mol L^{-1}, of the following gases in the container at 483 K.

        i. $H_2(g)$

        ii. $H_2S(g)$

    c) Calculate the value of the equilibrium constant, K_c, for the decomposition reaction at 483 K.

    d) Calculate the partial pressure of $S_2(g)$ in the container at equilibrium at 483 K.

    e) For the reaction $H_2(g)$ + _ $S_2(g)$ Δ $H_2S(g)$ at 483 K, calculate the value of the equilibrium constant, K_c.

AP Chemistry Review Key

Ch. 4

p.180

1. (#11) b. $MgCl_2 \rightarrow Mg_2^{+} + 2Cl^{-}$

d. $(NH_4)_2SO_4 \rightarrow 2NH_4^{+} + SO_4^{2-}$

2. (#17) d. M $Al^{3+} = 4.0 \text{ X } 10^{-3}$M, M $SO_4^{2-} = 6.0 \text{ X } 10^{-3}$M

3. (#29) a. $BaSO_4(s)$ d. $Fe(OH)_3(s)$

p. 182

4. (#43) a. $2KOH(aq) + Mg(NO_3)_2(aq) \rightarrow Mg(OH)_2(s) + 2KNO_3(aq)$

b. Mg(OH)2(s)

c. 0.583g

d. 0.100M K+, 0.200M NO3-, 5.00 X 10-2M Mg^{2+}

5. (#45) a. $2HClO_4(aq) + Mg(OH)_2(s) \gamma Mg(ClO_4)_2(aq) + 2H_2O(l)$

6. (#57) j. Na, +1; O, -2; C, +3

7. (#63) c. $7H_2O + 4H+ + 3As_2O_3 + 4NO_3^{-} \gamma 4NO + 6H_3AsO_4$

8. (#65) c. $OH- + H_2O + NO_2^{-} + Al \gamma NH_3 + 2AlO_2^{-}$

Ch. 7

Pg. 340

9. (#31) As electrons are removed, the nuclear charge exerted on the reaming electrons increase. Since the reaming electrons are "held" more strongly by the nucleus, the energy required to remove these electrons increases.

10. (#35) Yes, the maximum number of unpaired electrons in any configuration corresponds to a minimum in electron-electron repulsions.

11. (#55) n=1, 91.20 nm; n=2, 364.8 nm

P. 341

12. (59) n=1, 2, 3…; l=0,1,2….(n-1); m_l= -l, …, -2, -1, 0,1,2…,+l

13. (#61) b, l must be less than n; d, for l=0, m_l=0 only

14. (#87) a, Li

b, P

c, O^+

15. (#93) Se is an exception to the general ionization trend. There are extra electron-electron repulsions in Se because 2 electrons are in the same 4p orbital resulting in a lower ionization energy than expected.

Ch. 16

P. 829

16. (#25) a. –

b. +

17. (#31) a. Yes (delta G less than 0)

b. 196K

18. (#39) 262J/K * Mol

19. (#49) –300kJ

Ch.5

P. 233

20. (#35) 0.449 M

P. 234

21. (#55) a. $2CH_4(g) + 2NH_3(g) + 3O_2(g)$ γ $2HCN(g) + 6H_2O(g)$

b. 13.3 liters

22. (#65) a. X CH_4= .412, X O_2=0.588

b. 0.161 M

c. 1.06 g of CH_4, 3.03 g of O_2

P. 235

23. (#101) 24 torr

Ch. 8&9

P. 406

24. (#35) a. $Cu > Cu^{+} > Cu^{2+}$

b. $Pt^{2+} > Pd^{2+} > Ni^{2+}$

25. (#39) a. NaCl, Na^{+} smaller than K^{+}

e. Na_2O, O^{2-} greater charge than Cl^{-}

P. 407

26. (#47) a. –183kJ

b. –109kJ

27. (#55) 146 kJ/mol

P. 408

28. (#63)

29. (#79)

P. 410

30. (#85) a. linear, 180°

b. T- Shaped, ~90°

c. see-saw, ~90°

d. trigonal bipyramid, 90° and 120°

P. 444

31. (#23)

Chapter 10

P.504

32. (#35) a. London Dispersion (LD)

b. dipole, LD

c. hydrogen bonding, LD

d. ionic

e. LD

f. dipole, Ld

g. ionic

33. (#81) 89°C

34. (#87) 1680 kJ

P. 552

35. (#9) 0.365 *M*

36. (#13) 0.16 L

P. 553

37. (#27) 50.0%, 4.69 *M*, 5.49 *m*, 9.68 X 10^{-2}

38. (#35) KCL $\gamma\, k^{+}(aq)$ +$Cl^{-}(aq)$ ΔH_{soln} = 31kJ/mol

P. 554

39. (#41) Ability to form hydrogen bonding interactions, ability to break up into ions, and polarity are some factors affecting solute solubility.

    a. CH_3CH_2OH

    b. $CHCl_3$

    c. CH_3CH_2OH

40. (#47) 23.6 torr

P. 555

41. (#59) 100.042°C

42. (#61) 14.8 g

43. (#67) a. ΔT = 2.0 X 10-5°C, *pie* = 0.20 torr

b. Osmotic pressure is better for determining the molar mass of large molecules. A temperature change of 10^{-5}°C is very difficult to measure. A change in height of a column of mercury by 0.2mm is not as hard to measure precisely.

P. 556

44. (#71) a. 0.010 *m* Na_3PO_4 and 0.020 *m* KCl

b. 0.020 *m* HF

c. 0.020 *m* $CaBr_2$

Ch. 12

P. 602

45. (#11) a. The greater frequency of collisions, the greater is the opportunity for molecules to react, and hence the greater is the rate.

b. Chemical reactions involve the making and breaking of chemical bonds. The kinetic energy of the collisions can be used to break bonds. So, as the kinetic energy of the collisions increased, the rate increases.

c. For a reaction to occur, it is the reactive portion of each molecule that must be involved in a collision. Only some of all possible collisions have the correct orientation to convert from reactants to products.

46. (#23) a. rate = $k[NO]^2[Cl_2]$

b. 1.8 X $10^2 L^2/mol^2 \cdot$ min

P. 604

47. (#29) rate = $k[H_2O_2]$; $\ln[H_2O_2] = -kt + \ln[H_2O_2]_0$; $k = 8.3 \times 10^{-4} s^{-1}$; 0.037 *M*

P. 605

48. (#45) a. rate = $k[CH_3NC]$

b. rate = $k\,[O_3][NO]$

49. (#59) a. NO

b. NO_2

c. 2.3

Ch. 13

P. 650

50. (#19) b. $K = [NO_2]^2/[N_2O_4]$

51. (#21) a. 0.11

b. 77

P. 651

52. (#29) 1.5 X 10^8

53. (#31) b. $K = [N_2O][H_2O]^2$

54. (#37) a. $Q > K$, so reaction shifts left to reach equilibrium

b. $Q = K$, so reaction is at equilibrium.

c. $Q < K$, so reaction shifts right to reach equilibrium.

P. 652 & 653

55. (#45) 0.056

56. (#57) 4.8

57. (#63) a. left

b. right

c. left

d. no effect

e. no effect

f. right

Ch. 14

P. 712

58. (#25) a. H20 and CH3CO2--

b. CH3CO2--, since the equilibrium favors acetic acid ($K_a < 1$)

c. Acetate ion is a better base than water and produces basic solutions in water. When we put acetate ion unto water solution as the only major basic species, the reactions

$$CH_3CO_2^{--} + H_2O \Delta CH_3CO_2H + OH^{—}$$

Now the competition is between $CH_3CO_2^{--}$ and $OH^{—}$ for the proton. Hydroxide ion is the strongest base possible in water. The above equilibrium favors acetate plus water resulting in a K_b value less than one. Those species we specifically call weak bases ($10^{-14} < K_b < 1$) lie between H_2O and OH^- in base strength. Weak bases are stronger bases than water but are weaker bases than OH^-.

P. 713

59. (#57) [HOCl] = 1.5 *M*; $[H^+] = [OCl^-] = 2.3 \times 10^{-4}$ *M*; $[OH^-] = 4.3 \times 10^{-11}$ *M*; pH = 3.64

60. (#65) b. 1.9%

61. (#87) NH_3 and H_2O, 1.6×10^{-3} *M*, 11.20

62. (#89) a. $[OH^-] = 8.9 \times 10^{-3}$ M

Ch. 15

P. 783

63. (#53) a. 1.0 mol b. 0.30mol c. 1.3mol

64. (#65) a. 3.19, 8.45 b. 10.74, 5.96

65. (#69) a. yellow b. 8.0 c. blue

66. (#93) 2.3×10^{-11}mol/L

67. (#97) yes; $Q = 1.9 \times 10^{-4} > K_{sp}$

Ch. 17

P. 879

68. (#14) i.

69. (#15) Reactions a, b, and c are oxidation-reduction reactions.

	Oxidizing Agent	Reducing Agent	Substance Oxidized	Substance Reduced
a	H_2O	CH_4	CH_4(C)	H_2O(H)
b	$AgNO_3$	Cu	Cu	$AgNO_3$(Ag)
c	HCl	Zn	Zn	HCl(H)

70. (#35) a. $E^\circ = 0.46V$, spontaneous; b. $E^\circ = -0.53V$, not spontaneous;

71. (#45) 77 kJ/mol

72. (#55) a. larger b. smaller 73. (# 89) 1.14×10^{-2} M