SCH4U1 Exam Review

1. Name all the aldehyde isomers that have the molecular formula $C_5H_{10}O$. Name and draw all the isomers of $C_5H_{10.}$

2. Complete the following table by identifying an eight-carbon compound in each group.

Group	One structural formula	Name	Saturated/unsaturated
alkyne			
cycloalkane			
alcohol			
ether			

3. Butane can form a primary alcohol, a secondary alcohol, and a tertiary alcohol. Write the structural formula of each alcohol.

4. State the steps you would take to determine whether or not a molecule has an overall molecular polarity. Explain the significance of each step, and give an example.

5. What is the difference between addition polymerization and condensation polymerization? Give an example of a polymer formed by each reaction.

6. Describe and explain the general trends in atomic radius in the periodic table.

7. What are the four quantum numbers in the quantum mechanical model of the atom? What do these numbers represent?

8. Define the term "orbital." Describe the shape of each type of orbital. State how many of each type are present, and in which energy levels they are located.

9. What are the allowed values of m_l for an electron with each orbital-shape quantum number.
a) $l = 3$ b) $l = 1$

10. What are the allowed values of l for an electron with each principal quantum number.
a) $n = 4$ b) $n = 6$

11. Distinguish between the terms "ionization energy" and "electron affinity."

12. What is the relationship between atomic radius and ionization energy? Explain your answer.

13. What are the allowed values of l and m_l if $n = 2$? What is the total number of orbitals in this energy level?

14. What are the possible values of m_l if $n = 4$ and $l = 2$? What kind of orbital is described by these quantum numbers? How many orbitals can be described by these quantum numbers?

15. Using water as an example, differentiate between intermolecular forces and intramolecular forces.

16. List the different types of intermolecular forces, and give an example for each.

17. CF_4 contains polar C—F bonds. How do we know the bonds are polar? Is this molecule polar or non-polar? Identify the different shapes and their polarities.

18. The neutralization of nitric acid with potassium hydroxide has an enthalpy change of –53.4 kJ/mol. Write a thermochemical equation for this reaction.

19. The specific heat capacity of diamond is 0.5050 J/g•°C. How much energy is required to heat 25.0 g of diamond from 10.5°C to 15.6°C?

20. Methane burns in oxygen to form carbon dioxide and water. This process releases 882 kJ/mol of methane.
a) Write the thermochemical equation for this reaction.
b) If 15.0 g of methane is burned, how much heat is released?

21. What is the order of the reaction with the rate law expression $r = k[A][B]^2$?

22. What physical properties can be measured in each reaction?
a) $CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$
b) $2H_2O_{(g)} \rightarrow 2H_{2(g)} + O_{2(g)}$
c) $HCl_{(aq)} + Zn_{(s)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$
d) $N_2O_{5(g)} \rightarrow 2NO_{2(g)} + ½\, O_2$

23. The rate constant for the following reaction is $6.0 \times 10^{-4}\ s^{-1}$. What is the half-life of this reaction?

24. What is the difference between a homogeneous catalyst and a heterogeneous catalyst? Give an example of each type of catalyst.

26. Give an example of a chemical or physical process that results in an increase in entropy. Explain why this process is an example of an increase in entropy.

27. What is the equilibrium constant expression for the following equilibrium reaction?
$H_2O_{(g)} + Cl_2O_{(g)} \Leftrightarrow 2HOCl_{(g)}$

28. The equilibrium constant for the following reaction is 4.8×10^{-3} at 25ºC.
$N_2O_{4(g)} \Leftrightarrow 2NO_{2(g)}$
a) Which direction does the equilibrium normally favour?
b) Which species has the greater concentration?
c) What is the equilibrium constant for the reverse reaction?

29. List three different strategies that could be used to shift the following equilibrium toward the products.
$2NOCl_{(g)} \Leftrightarrow 2NO_{(g)} + Cl_{2(g)}$

30. List three different strategies that could be used to shift the following equilibrium toward the reactants. **Note:** List changing concentration only once.
$2SO_{3(g)} + 197\ kJ \Leftrightarrow 2SO_{2(g)} + O_{2(g)}$

31. Label the conjugate acid-base pairs in each reaction.
a) $CO_32-_{(aq)} + H_2O_{(l)} \Leftrightarrow HCO_3-_{(aq)} + OH^-_{(aq)}$
b) $H_2SO_{4(aq)} + H_2O_{(l)} \Leftrightarrow H_3O^+_{(aq)} + HSO_4-_{(aq)}$

32. Write the equilibrium expression for each equilibrium.
a) $HNO_{2(aq)} + H_2O_{(l)} \Leftrightarrow H_3O^+_{(aq)} + NO_2-_{(aq)}$
b) $HCOOH_{(aq)} + H_2O_{(l)} \Leftrightarrow HCOO^-_{(aq)} + H_3O^+_{(aq)}$

33. Calculate the pOH of a 0.025 mol/L solution of perchloric acid, $HClO_{3(aq).}$

34. Calculate the pH of a 0.000 25 mol/L solution of sodium hydroxide.

35. List the following acids in order, from strongest to weakest.

HCNO	$K_a = 3.5 \times 10^{-4}$
HF	$K_a = 6.8 \times 10^{-4}$
HIO	$K_a = 2.3 \times 10^{-11}$
HIO_3	$K_a = 1.69 \times 10^{-1}$

36. Which solution represents an equilibrium system? Explain your answer, based on the behaviour of the solute.
A. a solution of copper(II) sulfate, in which some of the water has evaporated and crystals of copper(II) sulfate are forming in solution
B. a teaspoon of sugar dissolved in a cup of hot tea cooling on a table
C. a supersaturated solution of sodium acetate at constant temperature

37. Write the equilibrium reaction and the equilibrium expression for the dissolving of silver sulfide.

38. Identify the oxidizing agent and the reducing agent in the following redox reaction. Explain your answers.
$MnO_4{-} + SO_3{2-} \rightarrow Mn^{+2} + SO_4{2-}$

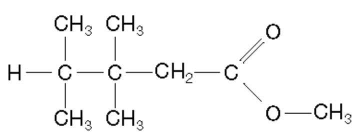
39. Identify the oxidizing agent and the reducing agent in the following redox reaction. Explain your answers.
$BrO_3{-} + I^- + H^+ \rightarrow Br^- + I_2 + H_2O$

40. Determine whether the following reaction is a redox reaction. Explain your answer.
$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$

41. Determine whether the following reaction is a redox reaction. Explain your answer.
$MnO_4{-} + CN^- \rightarrow MnO_2 + CNO^-$

42. What is the oxidation number of chlorine in each anion?
a) OCl^- b) $ClO_4{-}$

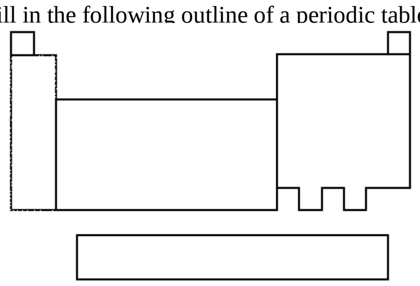
43. Complete the following table.

Name	Complete structural diagram	Class
2-fluoro-3,3-dimethylbutanal		
		
4-methylpentanamide		
metadimethylbenzene		

44. Draw the structural formula for each of the four isomers of butyl chloride (C_4H_9Cl). Name each isomer.

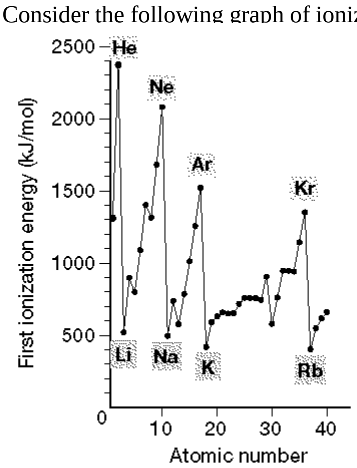
45. Use structural formulas to write the equation for the reaction of acetic(ethanoic) acid with 1-butanol.

46. Fill in the following outline of a periodic table to show the four energy sublevel (*s*, *p*, *d*, and *f*) blocks.



47. On the above outline of a periodic table, use arrows to indicate the directions of increasing
a) atomic radius b) ionization energy c) electron affinity.

48. Consider the following graph of ionization energy.



a) Describe the trend you observe down a group. Explain this trend, using the atomic model.
b) Describe the trend you observe across a group. Explain this trend, using the atomic model.

49. Complete the following table to describe the three intramolecular forces.

Force	Model/diagram	Nature of attraction	Energy (kJ/mol)	Example
ionic				
covalent				
metallic				

50. Create a table to summarize the differences in the physical properties of the different types of solids (atomic, molecular, covalent network, ionic, and metallic).

51. Determine the molecular shape of each molecule, and draw the three-dimensional representation.
a) $AsCl_5$ b) H_2CO c) OCl_2

52. When benzene is reacted with liquid bromine, bromobenzene is formed.
a) Write the balanced chemical equation for this reaction.
b) What type of reaction is it?
c) What amount of benzene is needed to react with excess bromine to produce 25 g of bromobenzene? Assume 100% yield.

53. Explain what is wrong with each set of quantum numbers.
a) $n = 3$, $l = 3$, $m_l = -2$; name: 3*d* b) $n = 5$, $l = 3$, $m_l = 5$; name: 5*f*

54. Explain what is wrong with each set of quantum numbers.
a) $n = 2$, $l = -1$, $m_l = 0$; name: 2*p* b) $n = 3$, $l = 2$, $m_l = -2$; name: 3*d*

55. Fill in the missing value(s) in each set of quantum numbers.
a) $n = ?$, $l = 1$, $m_l = -1$; name: 3*p* b) $n = 1$, $l = ?$, $m_l = ?$; name: 1*s*

56. Fill in the missing values in each set of quantum numbers.
a) $n = 3$, $l = 2$, $m_l = ?$; name: ? b) $n = ?$, $l = ?$, $m_l = -3$; name: 4*f*

57. What is the name of the orbital that is represented by each set of quantum numbers?
a) $n = 2$, $l = 1$, $m_l = -1$ b) $n = 4$, $l = 2$, $m_l = 0$

58. What is the name of the orbital that is represented by each set of quantum numbers?
a) $n = 1$, $l = 0$, $m_l = 0$ b) $n = 5$, $l = 3$, $m_l = -3$

59. Write the electron configuration for each element or ion.
a) Mn b) Ca c) Ag^+

60. Write the condensed electron configuration for each element or ion.
a) Bi b) Zn^{2+} c) Al

61. Draw the orbital diagram for each element or ion.
a) Fe b) Na^+ c) Cl

62. What types of intermolecular forces affect each compound or mixture? Explain your reasoning.
a) CH_2Cl_2 b) C_5H_{12}
c) O_2 dissolved in H_2O d) NaCl dissolved in H_2O

63. For each molecule; draw the Lewis structure, use the VSEPR theory to predict the shape of the compound, decide whether or not the molecule is polar, determine the bond angles and draw the dipole movement for the molecule if the molecule is polar
a) NH_4+ b) $AsCl_5$ c) ICl

64. For each molecule; draw the Lewis structure, predict the shape of the compound, decide whether or not the molecule is polar, determine the bond angles and draw the polarity for the molecule if the molecule is polar
a) CH_2Cl_2 b) PH_3 c) SeI_6

65. A 26.6 g sample of mercury is heated to 110.0°C and then placed in 125 g of water in a coffee-cup calorimeter. The initial temperature of the water is 23.00°C. The specific heat capacity of water is 4.184 J/g•°C, and the specific heat capacity of mercury is 0.139 J/g•°C. What is the final temperature of the water and the mercury?

66. Given equations (1), (2), and (3), calculate the heat of reaction for equation (4).

(1) $\frac{1}{2}N_{2(g)} + \frac{3}{2}H_{2(g)} \rightarrow NH_{3(g)}$	$\Delta H°_f = -46.15$ kJ
(2) $\frac{1}{2}N_{2(g)} + O_{2(g)} \rightarrow NO_{2(g)}$	$\Delta H°_f = 33.81$ kJ
(3) $H_{2(g)} + \frac{1}{2}O_{2(g)} \rightarrow H_2O_{(g)}$	$\Delta H°_f = -241.6$ kJ
(4) $NH_{3(g)} + \frac{7}{4}O_{2(g)} \rightarrow NO_{2(g)} + \frac{3}{2}H_2O_{(g)}$	

67. Given equations (1) and (2), calculate the heat of reaction for equation (3).

(1) $\frac{1}{4}P_{4(s)} + \frac{3}{4}O_{2(g)} \rightarrow \frac{1}{4}P_4O_6$	$\Delta H° = -410$ kJ
(2) $P_4O_{6(s)} + 2O_{2(g)} \rightarrow P_4O_{10(s)}$	$\Delta H°_s = -1344$ kJ
(3) $P_4H_{10(s)} \rightarrow P_{4(s)} + 5O_{2(g)}$	

68. Use standard heats of formation to calculate the heat of reaction for the following equation.
$2H_2S_{(g)} + 3O_{2(g)} \rightarrow 2H_2O_{(l)} + 2SO_{2(g)}$

69. The heat of formation of $NaClO_{3(s)}$ is −360.1 kJ/mol. Use this heat of formation, as well as standard heats of formation, to determine the heat of reaction for the following equation

$$.NaClO_{3(s)} \rightarrow NaCl_{(s)} + \frac{3}{2}O_{2(g)}$$

70. The experimental data in the table below were collected for the following reaction of nitrogen monoxide and hydrogen. What is the rate law for this reaction?
$2NO_{(g)} + 2H_{2(g)} \rightarrow N_{2(g)} + 2H_2O_{(g)}$

Trial	Initial concentration (mol/L)		Initial rate of disappearance of NO (mol/L•s)
	[NO]	$[H_2]$	
1	0.10	0.10	1.23×10^{-3}
2	0.10	0.20	2.46×10^{-3}
3	0.20	0.10	4.92×10^{-3}

71. Determine the equilibrium constant for the following reaction, based on the equilibrium concentrations below.
$H_{2(g)} + I_{2(g)} \Leftrightarrow 2HI$
$[H_2] = 1.00$ mol/L $[I_2] = 1.00$ mol/L $[HI] = 7.1$ mol/L

72. 0.150 mol of SO_3 and 0.150 mol of NO are placed in a 1 L container and allowed to react as follows. At equilibrium, the concentration of both SO_2 and NO_2 is 0.0621 mol/L. What is the equilibrium constant?
$SO_{3(g)} + NO_{(g)} \Leftrightarrow SO_{2(g)} + NO_{2(g)}$

73. Suppose that 0.350 mol of A and 0.520 mol of B are placed a 1.50 L container and the following hypothetical equilibrium is established. If the equilibrium amount of C is 0.150 mol, what is the equilibrium constant for this reaction? $A_{(g)} + 2B_{(g)} \Leftrightarrow 3C_{(g)}$

74. The following equilibrium system has an equilibrium constant of 2.0×10^9. Find the equilibrium concentrations if 0.250 mol of both $H_{2(g)}$ and $Br_{2(g)}$ are added to a 5.00 L container.
$H_{2(g)} + Br_{2(g)} \Leftrightarrow 2HBr_{(g)}$

75. $N_{2(g)}$ and $O_{2(g)}$ can exist in equilibrium with $NO_{(g)}$, as shown below. The equilibrium constant at 25.0°C is 4.8×10^{-31}. If initially there are 1.25 mol and 0.50 mol of oxygen in a 1.00 L vessel, find the equilibrium concentrations of each species. $N_{2(g)} + O_{2(g)} \Leftrightarrow 2NO_{(g)}$

76. The equilibrium constant for the following reaction is 0.18 at a set temperature. Find the equilibrium concentrations if the initial concentration of PCl_3 is 0.225 mol/L and the initial concentration of Cl_2 is 0.150 mol/L.
$PCl_{3(g)} + Cl_{2(g)} \Leftrightarrow PCl_{5(g)}$

77. A solution of hydrocyanic acid has an initial concentration of 5.0×10^{-3} mol/L. What are the concentrations of the ions at equilibrium, if $K_a = 4.9 \times 10^{-10}$.

78. A solution of nitrous acid has a concentration of 0.65 mol/L. What is the pH of the solution when equilibrium is established, if $K_a = 4.5 \times 10^{-4}$?

79. 30.0 mL of a solution of a diprotic acid, oxalic acid ($C_2H_2O_4$), is titrated with 56 mL of a 0.050 mol/L solution of potassium hydroxide. What was the concentration of the oxalic acid?

80. The solubility of calcium fluoride is 1.6×10^{-2}g/L at 20°C. Determine K_{sp} for calcium fluoride.

81. Calculate the concentrations of iron ions and hydroxide ions in a solution of iron(II) hydroxide ($K_{sp} = 1.8 \times 10^{-15}$).

82. What is the molar solubility of barium fluoride, if K_{sp} for barium fluoride is 1.7×10^{-6}?

83. If 25 mL of 2.00×10^{-2} mol/L sodium hydroxide is added to 80 mL of 3.2×10^{-2} mol/L magnesium chloride, will a precipitate form? K_{sp} for magnesium hydroxide is 1.2×10^{-11}.

84. 50.0 mL of 0.0015 mol/L calcium chloride solution is added to 75.0 mL of 0.010 mol/L sodium sulfate solution ($K_{sp} = 2.0 \times 10^{-4}$). Does a precipitate of calcium sulfate form?

85. a) State the oxidation number of each element in the following redox reaction.
$S_4O_62-_{(aq)} + Cr^{2+}_{(aq)} \rightarrow Cr^{3+}_{(aq)} + S_2O_32-_{(aq)}$
b) Identify the reactant oxidized and the reactant reduced.

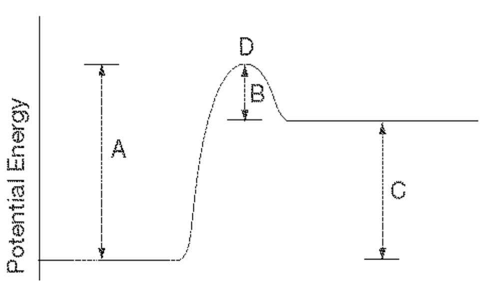
86. Balance the following redox reaction, using the half-reaction method under acidic conditions.
$NO_3- + Bi \rightarrow Bi^{3+} + NO_2$

87. Balance the following redox reaction using the half-reaction method under acidic conditions.
$Cr_2O_72- + I^- \rightarrow Cr^{3+} + I_2$

88. Balance the following redox reaction using the half-reaction method under basic conditions.
$ClO^- \rightarrow Cl^- + ClO_3-$

89. Balance the following redox reaction using the oxidation number method.
$MnO_{2(s)} + Al_{(s)} \rightarrow Mn_{(s)} + Al_2O_{3(s)}$

90. Use the following diagram to answer the questions below.



a) Is the reaction exothermic or endothermic? Explain.
b) What letter represents the activation energy of the forward reaction?
c) What letter represents the heat of reaction?
d) What letter represents the activation energy of the reverse reaction?

91. The forward activation energy of a reaction is 25 kJ/mol, and the heat of reaction is −286.4 kJ/mol.
a) Sketch a potential energy diagram for the reaction. Label the axes, the forward activation energy, the heat of reaction, the transition state, and the reactants and products.
b) Indicate the numerical values of the forward activation energy and the enthalpy change on your diagram.
c) Show and label the effect of a catalyst.

92. Use collision theory to explain how surface area affects reaction rate.

93. Use collision theory to explain how temperature affects reaction rate.

94. Use collision theory to explain how concentration affects reaction rate.

95. Use collision theory to explain how a catalyst affects reaction rate.

96. Complete the following table, based on the following equilibrium system.
$2Cl_{2(g)} + 2H_2O_{(g)} \Leftrightarrow 4HCl_{(g)} + O_{2(g)}$ $\Delta H = +133$ kJ

Stress	Equilibrium shift
increase in temperature	
increase in hydrogen chloride	
increase in pressure at constant volume	
decrease in volume	
addition of inert gas at constant pressure	

97. K_a for vitamin C (also known as asborbic acid, $C_6H_8O_6$) is 8.0×10^{-5}. If 15 g of vitamin C is dissolved in 1.00 L of water, what is the pH of the solution at equilibrium?

98. HCl is used to recover oil in oil wells by dissolving rocks, so the oil will flow more easily. In one process, a 10% by mass HCl solution is injected into an oil well to dissolve the rocks. If the density of the acid solution is 1.073 g/mL, what is the pH of the solution?

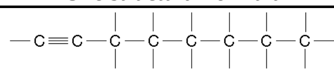
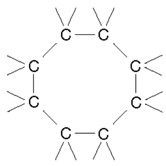
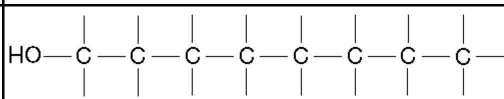
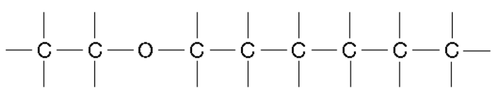
99. A student wants to determine K_{sp} for silver bromate. The student prepares a 1.0 L saturated solution of silver bromate at room temperature. The student then immerses a piece of copper in the solution and leaves it overnight. The next day, the student discovers that the copper weighs 0.26 g less than it did originally.
a) How many moles of copper reacted?
b) What is the balanced chemical equation for the reaction that occurred?
c) What is the concentration of silver ions in the original solution?
d) Determine K_{sp} for silver bromate at room temperature.

100. a) State the oxidation numbers of each element in the following redox reaction.
$Au_{(s)} + NO_3-_{(aq)} + 4H^+_{(aq)} \rightarrow Au^{3+}_{(aq)} + NO_{(g)} + 2H_2O_{(l)}$
b) Identify the reactant oxidized and the reactant reduced.

SCH4U1 Exam Review
Answer Section

1. pentanal, 3-methylbutanal, 2,2-dimethylpropanal, 2-methylbutanal

2.

Group	One structural formula	Name	Saturated/ unsaturated
alkyne		octyne	unsaturated
cycloalkane		cyclooctane	saturated
alcohol		1-octanol	saturated
ether		ethoxyhexane	saturated

Other answers are possible for the alcohol and the ether.

3. primary: C_4H_9OH; secondary: $C_2H_5CH(OH)CH_3$; tertiary: $CH_3C(CH_3)(OH)CH_3$

4. a) Determine whether or not polar bonds are present in the molecule by looking at the electronegativity differences (the ability to attract electrons) between the atoms within the molecule. High electronegativity differences (very polar bonds) between the atoms result in a bond dipole: a partially negative charge and a partially positive charge, separated by the length of the bond. An alcohol is a good example of a polar molecule, with the OH bond establishing regions that have partially negative and partially positive charges.
b) Consider the shape of the molecule. Is there a region with a partial negative charge and a region with a partial positive charge, not balanced by other, similar regions? Polar bonds exist between carbon and chlorine in a carbon tetrachloride molecule, but the shape is tetrahedral. This shape minimizes the electron repulsion forces within the molecule, so one region of the molecule does not appear more negative than any other region.

5. Addition polymerization is a reaction in which monomers with double bonds are joined together through a series of addition reactions. For example, polyethylene is formed from ethene. Condensation polymerization is a reaction in which monomers are joined together by the formation of ester or amide bonds and water is released. Nylon-66 is an example of a condensation polymer.

6. Atomic radius increases as you go down a group of elements in the periodic table. This trend is a result of increasing numbers of electrons occupying increasing numbers of energy levels. The effective nuclear charge changes only slightly and therefore does not offset the increase in size due to the increase in energy levels.
Atomic radius decreases as you go left to right across a period in the periodic table. The valence electrons are found in orbitals of the same energy level. At the same time, the effective nuclear charge is increasing with the increase in nuclear charge, which results in a greater force of attraction pulling the valence electrons closer to the nucleus. Thus, atomic size decreases.

7. 1) The principal quantum number, n, indicates the energy level of an atomic orbital and its relative size.
2) The orbital-shape quantum number, l, indicates the shape of the orbital.
3) The magnetic quantum number, m_l, indicates the orientation of the orbital.
4) The spin quantum number, m_s, indicates the direction in which the electron is spinning.

8. 1) The *s* orbital is spherical in shape. One *s* orbital is present in each energy level.
2) The three *p* orbitals are shaped like dumb-bells. They are present in all energy levels, starting with the second energy level.
3) There are five *d* orbitals. Four of them have four lobes arranged at right angles to each other, in one plane. One *d* orbital has two lobes and a ring of probability in the plane, at right angles to the lobes. The *d* orbitals are present in all energy levels higher than the third energy level.
4) The seven *f* orbitals have eight lobes arranged in three dimensions. They are present in all energy levels after the fourth energy level.

9. a) $m_l = -3, -2, -1, 0\ 1, 2, 3$ b) $m_l = -1, 0, 1$

10. a) $l = 0, 1, 2, 3$ b) $l = 0, 1, 2, 3, 4, 5$

11. Ionization energy is the energy that is required to completely remove one electron from a ground state gaseous atom. Electron affinity is the change in energy that occurs when an electron is added to a gaseous atom.

12. Ionization energy is the energy that is required to remove an electron completely from a ground state gaseous atom. This energy tends to increase as the atomic radius decreases. The closer the electrons are to the nucleus, the great the force of attraction pulling or holding the electrons in the atom.

13. If $n = 2$, then $l = 0, 1$. For $l = 0$, $m_l = 0$. For $l = 1$, $m_l = -1, 0, 1$. There are four orbitals in this energy level.

14. The values of m_l are $-2, -1, 0, 1, 2$. The five orbitals described by $l = 2$ are called the *d* orbitals.

15. Intermolecular forces are forces between molecules. Water has hydrogen bonds, dipole-dipole forces, and dispersion forces that hold adjacent water molecules together. Intramolecular forces are forces within a molecule. They hold the atoms together within the molecule. Each water molecule has polar covalent bonds between the hydrogen atoms and the oxygen atom.

16.

Intermolecular forces	Example
ion-dipole	sodium ions in water
hydrogen bonds	water
dipole-dipole	iodine monochloride
ion-induced dipole	ferrous ions and oxygen molecules
dipole-induced dipole	hydrochloric acid and chlorine
dispersion (London) forces	fluorine gas

17. The bond between C and F is a polar covalent bond because the electronegativity difference is greater than 0.4 but less than 1.7. Since CF_4 is tetrahedral in shape the effects of the bond dipoles cancel each other, resulting in a non-polar molecule.

18. Since the enthalpy change is negative, the reaction is exothermic. Therefore, the energy term is written on the product side of the reaction equation.

$$HNO_{3(aq)} + KOH_{(aq)} \rightarrow KNO_{3(aq)} + H_{2(l)} + 53.4 \text{ kJ}$$

19.

$\Delta H = mc\Delta T$
$= 25.0\ g \times 0.5050\ J/g\bullet°C \times 5.1°C$
$= 64.4\ J\frac{3}{2}$

20. a) $CH_{4(g)} + \frac{3}{2}O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)} + 882\ kJ$

b) $\frac{15.0\ g}{16.05\ g/mol} = 0.935\ mol$

$$\frac{882\ kJ}{1\ mol} = \frac{x\ kJ}{0.935\ mol}$$
$$x = 825\ kJ$$

Burning 15.0 g of methane releases 825 kJ of heat.

21. This is a third-order reaction. (The sum of the exponents is 1 + 2 = 3.)

22. a) mass of calcium carbonate, pH, volume of carbon dioxide gas if pressure constant, change in pressure if volume constant, change in electrical conductivity
b) change in pressure if constant volume, change in volume if constant pressure
c) pH, mass of zinc, volume of hydrogen at constant pressure or pressure due to hydrogen at constant volume, change in electrical conductivity

23. The equation has only $N_2O_{5(g)}$ as a reactant, and the molar coefficient is one. Therefore, this is a first-order reaction.
For a first-order reaction,

$t_{\frac{1}{2}} = 0.693 \div k$
$= 0.693 \div (6.0 \times 10^{-4}\ s^{-1})$
$= 1.16 \times 10^3\ s$
$= 19.3\ min$

The half-life is 19.3 min.

24. A homogeneous catalyst exists in the same phase as the reactants. An example is the use of $ZnCl_{2(aq)}$ as the catalyst in the reaction between aqueous solutions of $(CH_3)_2CHOH$ and HCl. A heterogeneous catalyst exists in a phase that is different from the phase of the reaction it catalyses. An example is the use of platinum or palladium to catalyse the hydrogenation of alkenes.

25. Favourable enthalpy changes have a negative sign, and favourable entropy changes have a positive sign.

26. A variety of examples can be given. One example is water changing from ice to liquid to water vapour. The change in state from solid to liquid to gas represents a positive change in entropy because the liquid molecules are more disordered than the solid molecules, and the gas molecules are even more disordered than the liquid molecules. Dissolving an ionic solid in water to form ions is another positive change in entropy, since the number of particles increases.

27. $K_{eq} = \frac{[HOCl]^2}{[H_2O][Cl_2O]}$

28.
a) Since the equilibrium constant is less than 1, the equilibrium favours the reactants.
b) $N_2O_{4(g)}$ has the greater concentration.
c) The equilibrium constant for the reverse reaction is the inverse of the equilibrium constant for the forward reaction.

$$K_{eq\,(reverse)} = \frac{1}{K_{eq\,(forward)}}$$
$$= \frac{1}{4.8 \times 10^{-3}}$$
$$= 2.1 \times 10^2$$

29. 1) Increase the concentration of $NOCl_{(g)}$.
2) Decrease the concentration of either $NO_{(g)}$ or $Cl_{2(g)}$.

3) Decrease the pressure.

30. 1) Decrease the temperature. 2) Increase the concentration of either $SO_{2(g)}$ or $O_{2(g)}$. 3) Increase the pressure.

31.

a)	$CO_3^{2-}{}_{(aq)}$	+	$H_2O_{(l)}$	⇔	$HCO_3^{-}{}_{(aq)}$	+	$OH^{-}{}_{(aq)}$
	base 1		acid 2		conjugate acid 1		conjugate base 2
b)	$H_2SO_{4(aq)}$	+	$H_2O_{(l)}$	⇔	$H_3O^{+}{}_{(aq)}$	+	$HSO_4^{-}{}_{(aq)}$
	acid 1		base 2		conjugate acid 2		conjugate base 1

$$\frac{[H_3O^+{}_{(aq)}][NO_2{}^-{}_{(aq)}]}{[HNO_{2(aq)}]}$$

$$\frac{[H_3O^+{}_{(aq)}][HCOO^-{}_{(aq)}]}{[HCOOH_{(aq)}]}$$

b) $K_a =$

33.
$pH = -\log[H_3O^+]$
$= -\log[0.025] = 1.60$
$pOH = 14 - pH$
$= 14 - 1.60 = 12.40$

34.
$pOH = -\log[OH^-]$
$= -\log(0.00025) = 3.60$
$pH = 14 - pOH$
$= 14 - 3.60 = 10.40$

35. HIO_3, HF, HCNO, HIO

36. Only solution A represents an equilibrium system. Equilibrium exists between the undissolved crystals and the dissolved crystals. In solution B, the temperature is changing. Therefore, there is not a condition of constant macroscopic properties. In solution C, all the solute has dissolved. Therefore, there is no solid that could be dissolving at the same rate as the solute is precipitating from the solution.

37. $Ag_2S_{(s)} \Leftrightarrow 2Ag^+{}_{(aq)} + S^{2-}{}_{(aq)}$
$K_{sp} = [Ag^+{}_{(aq)}]^2[S^{2-}{}_{(aq)}]$

38. The oxidizing agent is MnO_4^- because it gained electrons from SO_3^{2-} to allow SO_3^{2-} to be oxidized. The reducing agent is SO_3^{2-} because it lost electrons to MnO_4^- to allow MnO_4^- to be reduced.

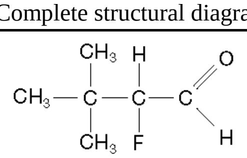
39. The oxidizing agent is BrO_3^- because it gained electrons from I^- to allow I^- to be oxidized. The reducing agent is I^- because it lost electrons to BrO_3^- to allow BrO_3^- to be reduced.

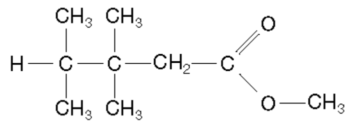
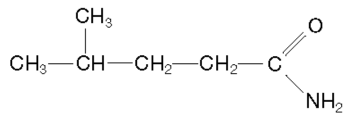
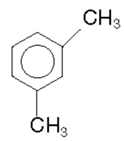
40. The reaction is not a redox reaction because there are no changes in the oxidation numbers of the elements.
$\underset{+6\ -2}{SO_{3(g)}} + \underset{+1\ -2}{H_2O_{(l)}} \rightarrow \underset{+1\ +6\ -2}{H_2SO_{4(aq)}}$

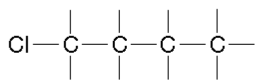
41.
The reaction is a redox reaction because there are changes in the oxidation numbers of the reactants. MnO_4^- is reduced, and CN^- is oxidized.
$\underset{+7\ -2}{MnO_4^-} + \underset{+2\ -3}{CN^-} \rightarrow \underset{+4\ -2}{MnO_2} + \underset{+4\ -3\ -2}{CNO^-}$

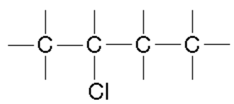
42. a) +1 b) +7

43.

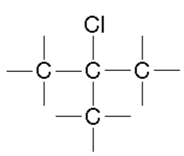
Name	Complete structural diagram	Class
2-fluoro-3,3-dimethylbutanal		aldehyde

methyl-3,3,4-trimethylpentanoate		ester
4-methylpentanamide		amide
metadimethylbenzene		aromatic

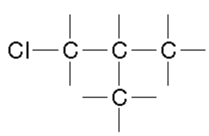
1-chlorobutane 

2-chlorobutane 

2-chloro-2-methylpropane

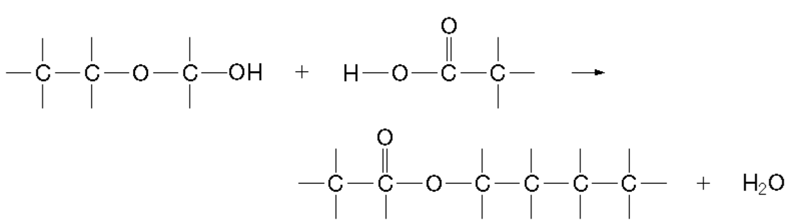


1-chloro-2-methylpropane

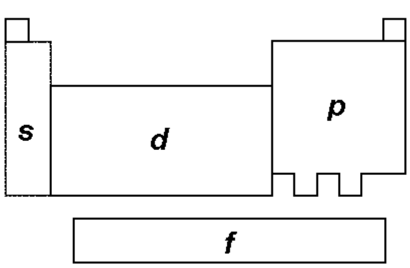


44.

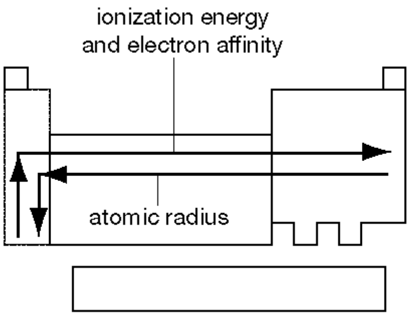
45.



46.

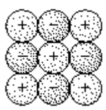
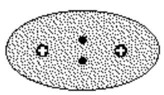
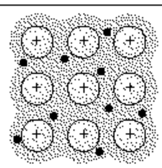


47.



48. a) Ionization energy decreases as you go down a group. This is caused by electrons occupying higher energy levels and therefore being farther from the nucleus. The increase in distance makes it easier to remove an electron completely from an atom. Thus, ionization energy decreases.
b) As you go left to right across a period, ionization energy increases. This is caused by increasing effective nuclear charge, resulting from increasing nuclear charge from left to right.

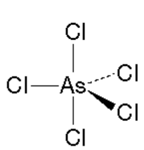
49.

Force	Model	Nature of Attraction	Energy (kJ/mol)	Example
ionic		cation-anion	400–4000	NaCl
covalent		nuclei-shared electron pair	150–1100	H—H
metallic		cations-delocalized electrons	75–1000	Fe

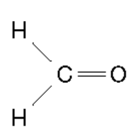
50. ANS:

Type of Crystal Solid	Boiling Point	Electrical Conductivity in Liquid State	Other Physical Properties of Crystals
Atom	low	very low	very soft
Molecular	generally low (non-polar); intermediate polar	very low	non-polar: very soft; soluble in non-polar solvents
			polar: somewhat hard, but brittle; many are soluble in water
Covalent Network	very high	low	hard crystals that are insoluble in most liquids
Ionic	high	high	hard and brittle; many dissolve in water
Metallic	most high	very high	all have a lustre, are malleable and ductile, and are good conductors; they dissolve in other metals to form alloys

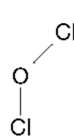
51. ANS:



a) trigonal bipyramidal



b) trigonal planar



c) bent

52.a) $C_6H_6 + Br_2 \rightarrow C_6H_5Br + HBr$

b) This is a substitution reaction.
c) 25 g of bromobenzene = 25 g ÷157 g/mol = 0.478 mol
This number of moles of benzene is needed, based on the balanced chemical equation.
Mass of 0.478 mol of benzene = 0.478 mol × 78 g/mol = 37.3 g

53.
a) l cannot equal 3, since the maximum value of l is $(n - 1)$ and d orbitals have an l value of 2.
b) m_l cannot equal 5, since the maximum value of m_l is $+l$.

54. ANS:
a) l cannot equal –1, since values of l can only be positive integers.
b) This set of quantum numbers is a valid set.

55. ANS:
a) $n = 3$ b) $l = 0$, $m_l = 0$

56. ANS:
a) $m_l = -2, -1, 0, 1,$ or 2; name: 3d b) $n = 4$, $l = 3$

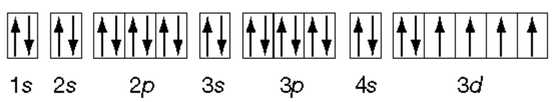
57. ANS:
a) 2p b) 4d

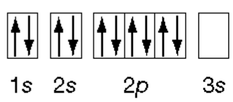
58. ANS:
a) 1s b) 5f

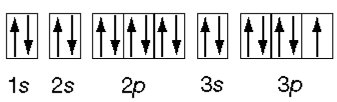
59. ANS:
a) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^5$ b) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$ c) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^8$

60. ANS:
a) $[Xe]\ 6s^2\ 4f^{14}\ 5d^1\ 6p^3$ b) $[Ar]\ 4s^2\ 3d^{10}$ c) $[Ne]\ 3s^2\ 3p^1$

61. ANS:

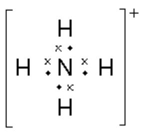
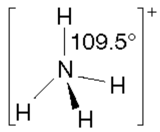
a) 

b) 

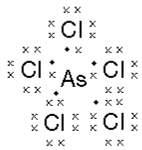
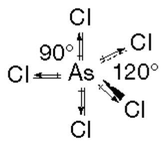
c) 

62. ANS:
a) Dipole-dipole: The molecules of CH_2Cl_2 have a molecular dipole.
b) Dispersion forces: The C_5H_{12} (hexane) molecule has covalent bonds between atoms.
c) Dipole-induced dipole: The H_2O molecule has a bent shape, so it has a molecular dipole. This dipole induces a dipole in the O_2 molecules, resulting in their mutual attraction.
d) Ion-dipole interactions: NaCl dissociates into its ions as it dissolves in H_2O. The H_2O molecule has a molecular dipole. Hence, ion-dipole interaction takes place between the ions of NaCl and the polar H_2O molecules.

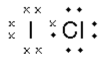
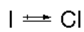
63. ANS:

a)   charged ion

tetrahedral

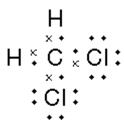
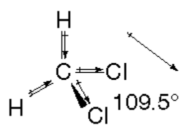
b)   non-polar molecule

trigonal bypyramidal

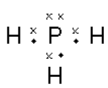
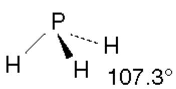
c)   polar molecule

linear

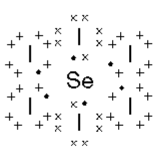
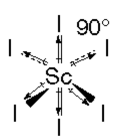
64. ANS:

a)   polar molecule

tetrahedral

b)   non-polar molecule

trigonal pyramidal

c)   non-polar molecule

octahedral

65. ANS:

$Q_{\text{lost by mercury}}$	=	$Q_{\text{gained by water}}$
$mc\Delta T$	=	$mc\Delta T$
(26.6 g)(0.139 J/g•°C)(110.0°C – T_f)	=	(125 g)(4.184 J/g•°C)(T_f – 23.00°C)
406.7 – 3.70T_f	=	523T_f – 12 029
519.3T_f	=	124 357
T_f	=	23.9°C

The final temperature of both the mercury and the water is 23.9°C.

66. ANS:

$NH_{3(g)} \rightarrow \frac{1}{2} N_{2(g)} + \frac{3}{2} H_{2(g)}$	$\Delta H°_f = 46.15$ kJ
$\frac{1}{2} N_{2(g)} + O_{2(g)} \rightarrow NO_{2(g)}$	$\Delta H°_f = 33.81$ kJ
$\frac{3}{2} H_{2(g)} +$	$\Delta H°_f = –362.4$ kJ

$\frac{3}{4}$ $O_{2(g)} \rightarrow$ $\frac{3}{2}$ $H_2O_{(g)}$	
$NH_{3(g)}$ + $\frac{7}{4}$ $O_{2(g)} \rightarrow NO_{2(g)}$ + $\frac{3}{2}$ $H_2O_{(g)}$	$\Delta H°_f = -281.9$ kJ

67. ANS:

$P_4O_{10(s)} \rightarrow P_4O_{6(s)} + 2O_{2(g)}$	$\Delta H° = +1344$ kJ
$P_4O_{6(s)} \rightarrow P_{4(s)} + 3O_{2(g)}$	$\Delta H° = +1640$ kJ
$P_4H_{10(s)} \rightarrow P_{4(s)} + 5O_{2(g)}$	$\Delta H° = +2984$ kJ

68. ANS:

$\Delta H_{rxn} = [2(\Delta H°_f\ SO_{2(g)} + 2(\Delta H°_f\ H_2O_{(l)})] - [2(\Delta H°_f\ H_2S_{(g)}) + 3(\Delta H°_f\ O_{2(g)})]$

$= [2(-296.8\text{ kJ/mol}) + 2(-285.8\text{ kJ/mol})] - [2(-20.6\text{ kJ/mol}) + 3(0.0\text{ kJ/mol})]$

$= -1124$ kJ/2 mol $H_2S_{(g)}$

$= -562$ kJ/mol $H_2S_{(g)}$

The heat of reaction is -562 kJ/mol $H_2S_{(g)}$.

69. ANS: $\frac{3}{2}$

$\Delta H_{rxn} = [\Delta H°_f\ NaCl_{(s)}$ + $\frac{3}{2}\Delta H°_f\ O_{2(g)})] - [\Delta H°_f\ NaClO_{3(s)}]$

$= [-411.2\text{ kJ/mol}) +$

$(0.0\text{ kJ/mol})] - [-360.1\text{kJ/mol})]$

$= -51.1$ kJ/mol $NaClO_{3(s)}$

The heat of reaction is -51.1 kJ/mol $C_2H_5OH_{(l)}$.

70. ANS:

Compare trials 1 and 2: Doubling the concentration of $H_{2(g)}$ causes the reaction rate to double. Therefore, the reaction rate is first order with respect to the concentration of $H_{2(g)}$.

Compare trials 1 and 3: Doubling the concentration of $NO_{(g)}$ causes the reaction rate to quadruple. Therefore, the reaction rate is second order with respect to $NO_{(g)}$.

React $[H_{2(g)}]^1[NO_{(g)}]^2$

71. ANS: $\frac{[HI]^2}{[H_2][I_2]}$

$K_{eq} = \frac{(7.1\text{ mol/L})^2}{(1.00\text{ mol/L})(1.00\text{ mol/L})}$

=

= 50.4

72. ANS:

Concentration (mol/L)	$SO_{3(g)}$	+	$NO_{(g)}$	⇔	$SO_{2(g)}$	+	$NO_{2(g)}$
Initial	0.150		0.150		0.0		0.0
Change	–0.0621		–0.0621		0.0621		0.0621
Final	0.0879		0.0879		0.0621		0.0621

$$\frac{[NO_2][SO_2]}{[NO][SO_3]}$$

$K_{eq} = \dfrac{(0.0621)(0.0621)}{(0.0879)(0.0879)}$

=

$= 4.99 \times 10^{-1}$

The equilibrium constant is 4.99×10^{-1}.

73. ANS:

Concentration (mol/L)	$A_{(g)}$	+	$2B_{(g)}$	⇔	$3C_{(g)}$
Initial	0.350 mol/1.5L		0.520 mol/1.5 L		0.0
	= 0.233		= 0.347		
Change	$-\frac{1}{3}(0.010)$ = −0.033		$-\frac{2}{3}(0.010)$ = −0.067		0.100
Final	0.200		0.280		0.150 mol/1.5 L = 0.100

$$\frac{[C]^3}{[A][B]^2}$$

$K_{eq} = \dfrac{(0.100)^3}{(0.200)(0.280)^2}$

=

$= 6.37 \times 10^{-2}$

The equilibrium constant is 6.37×10^{-2}.

74. ANS:

Concentration (mol/L)	$H_{2(g)}$	+	$Br_{2(g)}$	⇔	$2HBr_{(g)}$
Initial	0.250mol/5.00 L		0.250mol/5.00 L		0.0
	= 0.05		= 0.05		
Change	$-x$		$-x$		$2x$
Final	$0.05 - x$		$0.05 - x$		$2x$

$$\frac{[HBr]^2}{[H_2][Br_2]}$$

$K_{eq} = \dfrac{(2x)^2}{(0.05 - x)(0.05 - x)}$

=

$= 2.0 \times 10^9$

Take the square root of both sides.

$\dfrac{2x}{0.05 - x} =$	4.47×10^4
$2x =$	$(4.47 \times 10^4)(0.05 - x)$
$2x =$	$(2.235 \times 10^3) - (4.47 \times 10^4)x$
$2x + (4.47 \times 10^4)x =$	2.235×10^3
$x =$	0.05 mol/L

The equilibrium concentration of $HBr_{(g)}$ is 0.05 mol/L. The equilibrium concentrations of $H_{2(g)}$ and $Br_{2(g)}$ are both approximately zero.
Alternative approach: Since the equilibrium constant is so much greater than 1, assume that the reaction lies very much to the right. Hence, the reaction will have mainly products and very little reactants.

75. ANS:

Concentration (mol/L)	$N_{2(g)}$	+	$O_{2(g)}$	⇔	$2NO_{(g)}$
Initial	1.25		0.50		0.0
Change	$-x$		$-x$		$2x$
Final	$1.25 - x$		$0.50 - x$		$2x$

$$\frac{[NO]^2}{[O_2][N_2]}$$

$$\frac{(2x)^2}{(1.25-x)(0.50-x)}$$

=

$= 4.8 \times 10^{-31}$

Since the equilibrium constant is so small, the concentrations of the products will be extremely small. Therefore, $1.25 - x$ is approximately equal to 1.25, and $0.50 - x$ is approximately equal to 0.50. Since the expression is a perfect square, take the square root of both sides.

$\frac{(2x)^2}{(1.25)(0.50)} =$	4.8×10^{31}
$(2x)^2 =$	3.0×10^{31}
$2x =$	5.48×10^{16}
$x =$	2.74×10^{16}

The equilibrium concentration of $NO_{(g)}$ is $2(2.74 \times 10^{-16}) = 5.48 \times 10^{-16}$ mol/L. The equilibrium concentration of $N_{2(g)}$ is 1.25 mol/L, and the equilibrium concentration of $O_{2(g)}$ is 0.50 mol/L.

76. ANS:

Concentration (mol/L)	$PCl_{3(g)}$	+	$Cl_{2(g)}$	$\Leftrightarrow$	$PCl_{5(g)}$
Initial	0.225		0.150		0.0
Change	$-x$		$-x$		x
Final	$0.225 - x$		$0.150 - x$		x

$$K_{eq} = \frac{[PCl_5]}{[PCl_3][Cl_2]}$$

$$= \frac{x}{(0.225-x)(0.150-x)}$$

$$= 0.18$$

$$\frac{x}{0.03375 - 0.375x + x^2} = 0.18$$

$x = 0.18(0.033\ 75 - 0.375x + x^2)$

$= (6.075 \times 10^{-3}) - 0.0675x + 0.18x^2$

$0.18x^2 - 1.0675x + (6.075 \times 10^{-3}) = 0$

Solve for x using the general formula for a quadratic equation.

$x = 0.0057$ mol/L

The equilibrium concentration of $PCl_{5(g)}$ is 0.0057 mol/L, the equilibrium concentration of $PCl_{3(g)}$ is 0.219 mol/L, and the equilibrium concentration of $Cl_{2(g)}$ is 0.144 mol/L.

77. ANS:

Concentration (mol/L)	$HCN_{(aq)}$	+	$H_2O_{(l)}$	$\Leftrightarrow$	$H_3O^+{}_{(aq)}$	+	$CN^{-(aq)}$
Initial	5.0×10^{-3}				0.0		0.0
Change	$-x$				x		x
Final	$(5.0 \times 10^{-3}) - x$				x		x

$$\frac{[H_3O^+][CN^-]}{[HCN]}$$

$K_a = \frac{x^2}{(5.0 \times 10^{-3}) - x}$

=

$= 4.9 \times 10^{-10}$

Since x is very small, use an approximation.

$\frac{x^2}{5.0 \times 10^{-3}} =$	4.9×10^{-10}
$x^2 =$	2.45×10^{-12}
$x =$	1.6×10^{-6}

The equilibrium concentrations of $H_3O^+_{(aq)}$ and $CN^-_{(aq)}$ are 1.6×10^{-6} mol/L. The equilibrium concentration of $HCN_{(aq)}$ is 5.0×10^{-3} mol/L.

78. ANS:

Concentration (mol/L)	$HNO_{2(aq)}$	+	$H_2O_{(l)}$	⇔	$H_3O^+_{(aq)}$	+	$NO_2^-{}_{(aq)}$
Initial	0.65				0.0		0.0
Chan	$-x$				x		x
Final	$0.65 - x$				x		x

$$\frac{[H_3O^+][NO_2^-]}{[HNO_2]}$$

$K_a = \frac{x^2}{0.65 - x}$

=

$= 4.5 \times 10^{-4}$

$x^2 = (0.65 - x)(4.5 \times 10^{-4})$

$x^2 = (2.9 \times 10^{-4}) - (4.5 \times 10^{-4})x$

$x^2 + (4.5 \times 10^{-4})x - (2.9 \times 10^{-4}) = 0$

Use the general formula for the quadratic equation.

$x = 0.0168$

The equilibrium concentration of $H_3O^+_{(aq)}$ is 0.0168 mol/L.

$pH = -\log[H_3O^+]$

$= -\log(0.0168)$

$= 1.77$

The pH of the solution is 1.77.

79. ANS:

$2KOH_{(aq)} + C_2H_2O_{4(aq)} \Leftrightarrow K_2C_2O_{4(aq)} + 2H_2O_{(l)}$

Moles KOH = $C \times V$

$= 0.056\ L \times 0.050\ mol/L$

$= 2.8 \times 10^{-3}$ mol

$\frac{1\ mol\ C_2H_2O_4}{2\ mol\ KOH}$

Moles $C_2H_2O_4 = 2.8 \times 10^{-3}$ mol KOH ×

$= 1.4 \times 10^{-3}$ mol

$C_{oxalic\ acid} = \frac{n}{V}$ $\frac{10^{-3}\ mol}{0.030\ L}$

$= 1.4 \times$

$= 4.7 \times 10^{-2}$ mol/L

The concentration of the oxalic acid was 4.7×10^{-2} mol/L.

80. ANS:

$CaF_{2(s)} \Leftrightarrow Ca^{2+}_{(aq)} + 2F^-_{(aq)}$

K_{sp}	$= [Ca^{2+}_{(aq)}][F^-_{(aq)}]^2$
$[CaF_{2(s)}]$	$= \dfrac{1.6 \times 10^{-2}\ g/L}{78.01\ g/mol}$
	$= 2.05 \times 10^{-4}$ mol/L
$[Ca^{2+}_{(aq)}]$	$= [CaF_{2(s)}] = 2.05 \times 10^{-4}$ mol/L
$[F^-_{(aq)}]$	$= 2[CaF_{2(s)}]$
	$= 2(2.05 \times 10^{-4}$ mol/L)
	$= 4.10 \times 10^{-4}$ mol/L
K_{sp}	$= [Ca^{2+}_{(aq)}][F^-_{(aq)}]^2$
	$= (2.05 \times 10^{-4}$ mol/L)$(4.10 \times 10^{-4}$ mol/L$)^2$
	$= 3.45 \times 10^{-12}$

Therefore, K_{sp} for calcium fluoride is 3.45×10^{-12}.

81. ANS:
$Fe(OH)_{2(s)} \Leftrightarrow Fe^{2+}_{(aq)} + 2OH^-_{(aq)}$
$K_{sp} = [Fe^{2+}_{(aq)}][OH^-_{(aq)}]^2$

Concentration (mol/L)	$Fe(OH)_{2(s)}$ $\Leftrightarrow$	$Fe^{2+}_{(aq)}$ +	$2OH^-_{(aq)}$
Initial	----	0	0
Change	----	x	$2x$
Equilibrium	----	x	$2x$

K_{sp}	$= [Fe^{2+}_{(aq)}][OH^-_{(aq)}]^2$
K_{sp}	$= (x)(2x)^2$
1.8×10^{-15}	$= 4x^3$
x	$= 7.66 \times 10^{-6}$ mol/L
$[Fe^{2+}_{(aq)}]$	$= 7.66 \times 10^{-6}$ mol/L
$[OH^-_{(aq)}]$	$= 2(7.66 \times 10^{-6}$ mol/L)
	$= 1.53 \times 10^{-5}$ mol/L

82. ANS:
$BaF_{2(s)} \Leftrightarrow Ba^{2+}_{(aq)} + 2F^-_{(aq)}$
$K_{sp} = [Ba^{2+}_{(aq)}][F^-_{(aq)}]^2$

Concentration (mol/L)	$BaF_{2(s)} \Leftrightarrow$	$Ba^{2+}_{(aq)}$ +	$2F^-_{(aq)}$
Initial	----	0	0
Change	----	x	$2x$
Equilibrium	----	x	$2x$

$K_{sp} = [Ba^+_{(aq)}][F^-_{(aq)}]^2$
$K_{sp} = (x)(2x)^2$
$1.7 \times 10^{-6} = 4x^3$
$x = 7.5 \times 10^{-3}$ mol/L

The molar solubility of barium fluoride is 7.5×10^{-3} mol/L.

83. ANS:
$Mg(OH)_{2(s)} \Leftrightarrow Mg^{2+}_{(aq)} + 2$ $\dfrac{V_{initial}}{V_{final}}$

$[Mg^{2+}_{(aq)}] = [MgCl_{2(aq)}] \times$ $\dfrac{0.080\ L}{0.105\ L}$

$= (3.2 \times 10^{-2}$m $\dfrac{V_{initial}}{V_{final}}$

$= 2.4 \times 10^{-2}$ m

$[OH^-_{(aq)}] = [NaOH_{(aq)}] \times$ $\dfrac{0.025\ L}{0.105\ L}$

$= (2.00 \times 10^{-2}$ mol/L) $\times$

$= 4.8 \times 10^{-3}$ mol/L

$Q_{sp} = [Mg^{2+}_{(aq)}][OH^-_{(aq)}]^2$

$= (2.4 \times 10^{-2}\ mol/L)(4.8 \times 10^{-3}\ mol/L$ $\frac{V_{initial}}{V_{final}}$

$= 5.5 \times 10^{-7}$

Since $Q_{sp} > K_{sp}$, $Mg(OH)_2$ will precipitate until 5 $\times 10^{-7}$.

84. $[Ca^{2+}_{(aq)}]$ $\frac{0.050\ L}{0.125\ L}$ mol/L ×

$= 0.0015\ mol/L \times \frac{V_{initial}}{V_{final}}$ $= 6.0 \times 10^{-4}\ mol/L$

$[SO_42-_{(aq)}] = 0.010\ mol/L \times \frac{0.075\ L}{0.125\ L}$

$= 0.010\ mol/L \times$

$= 6.0 \times 10^{-3}\ mol/L$

$CaSO_{4(s)} \Leftrightarrow Ca^{2+}_{(aq)} + SO_42-_{(aq)}$

$K_{sp} = [Ca^{2+}_{(aq)}][SO_42-_{(aq)}]$

$Q_{sp} = [Ca^{2+}_{(aq)}][SO_42-_{(aq)}]$

$= (6.0 \times 10^{-4})(6.0 \times 10^{-3})$

$= 3.6 \times 10^{-6}$

Since $Q_{sp} < K_{sp}$, no precipitate forms.

85. ANS:

a) $\underset{+5/2\ \ -2}{S_4O_62-_{(aq)}} + \underset{+2}{Cr^{2+}_{(aq)}} \rightarrow \underset{+3}{Cr^{3+}_{(aq)}} + \underset{+2\ \ -2}{S_2O_32-_{(aq)}}$

b) The reactant oxidized is $Cr^{2+}_{(aq)}$. The reactant reduced is $S_4O_62-_{(aq)}$.

86. ANS:

Oxidation half-reaction:

$Bi \rightarrow Bi^{3+}$

$Bi \rightarrow Bi^{3+} + 3e^-$

Reduction half-reaction:

$NO_3- \rightarrow NO_2$

$NO_3- \rightarrow NO_2 + H_2O$

$2H^+ + e^- + NO_3- \rightarrow NO_2 + H_2O$

$2H^+ + NO_3- \rightarrow NO_2 + H_2O$

$2H^+ + e^- + NO_3- \rightarrow NO_2 + H_2O$

$6H^+ + 3e^- + 3NO_3- \rightarrow 3NO_2 + 3H_2O$

Total equation:

$6H^+ + 3e^- + 3NO_3- + Bi \rightarrow 3NO_2 + 3H_2O + Bi^{3+} + 3e^-$

Remove $3e^-$ to get the balanced redox equation:

$6H^+ + 3NO_3- + Bi \rightarrow 3NO_2 + 3H_2O + Bi^{3+}$

87. ANS:

Oxidation half-reaction:

$2I^- \rightarrow I_2$

$2I^- \rightarrow I_2 + 2e^-$

$6I^- \rightarrow 3I_2 + 6e^-$

Reduction half-reaction:

$Cr_2O_72- \rightarrow 2Cr^{3+}$

$Cr_2O_72- \rightarrow 2Cr^{3+} + 7H_2O$

$14H^+ + Cr_2O_72- \rightarrow 2Cr^{3+} + 7H_2O$

$14H^+ + 6e^- + Cr_2O_72- \rightarrow 2Cr^{3+} + 7H_2O$

Total equation:

$14H^+ + 6e^- + Cr_2O_72- + 6I^- \rightarrow 2Cr^{3+} + 7H_2O + 3I_2 + 6e^-$

Remove $6e^-$ to get the balanced redox equation:

$14H^+ + Cr_2O_72- + 6I^- \rightarrow 2Cr^{3+} + 7H_2O + 3I_2$

88. ANS:

Oxidation half-reaction:

$ClO^- \rightarrow ClO_3-$

$2H_2O + ClO^- \rightarrow ClO_3-$

$2H_2O + ClO^- \rightarrow 4H^+ + ClO_3-$

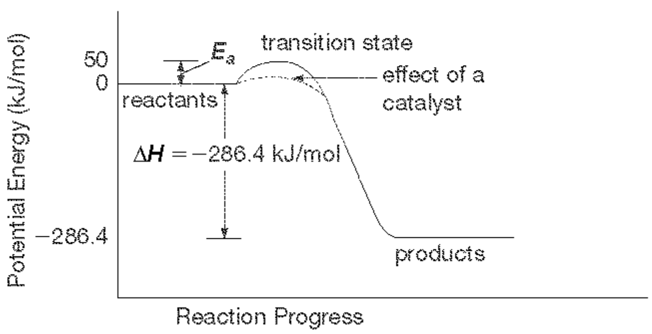
$4OH^- + 2H_2O + ClO^- \rightarrow ClO_3- + 4H^+ + 4OH^-$

$4OH^- + 2H_2O + ClO^- \rightarrow 4H_2O + ClO_3–$
$4OH^- + ClO^- \rightarrow 2H_2O + ClO_3–$
$4OH^- + ClO^- \rightarrow 2H_2O + 4e^- + ClO_3–$
Reduction half-reaction:
$ClO^- \rightarrow Cl^-$
$ClO^- \rightarrow Cl^- + H_2O$
$ClO^- + 2H^+ \rightarrow Cl^- + H_2O$
$ClO^- + 2H^+ + 2OH^- \rightarrow Cl^- + H_2O + 2OH^-$
$ClO^- + 2H_2O \rightarrow Cl^- + H_2O + 2OH^-$
$ClO^- + H_2O \rightarrow Cl^- + 2OH^-$
$ClO^- + 2e^- + H_2O \rightarrow Cl^- + 2OH^-$
$2ClO^- + 4e^- + 2H_2O \rightarrow 2Cl^- + 4OH^-$
Total equation:
$4OH^- + ClO^- + 2ClO^- + 4e^- + 2H_2O \rightarrow 2Cl^- + 4OH^- + 2H_2O + 4e^- + ClO_3–$
Remove $4e^-$, $2H_2O$, and $4OH^-$ to get the final equation:
$3ClO^- \rightarrow 2Cl^- + ClO_3–$

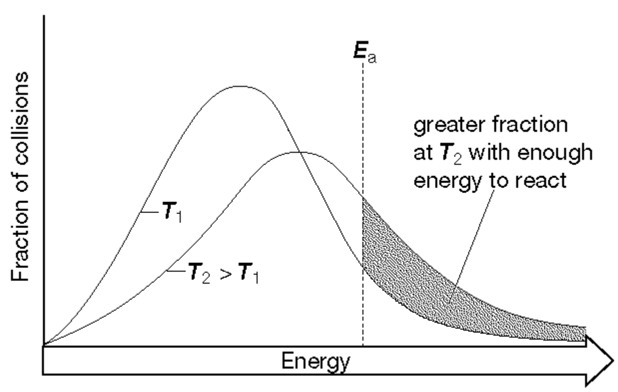
89. ANS:
$MnO_{2(s)} + Al_{(s)} \rightarrow Mn_{(s)} + Al_2O_{3(s)}$
+4 –2 0 0 +3 –2
$3MnO_{2(s)} + 4Al_{(s)} \rightarrow Mn_{(s)} + Al_2O_{3(s)}$
$3MnO_{2(s)} + 4Al_{(s)} \rightarrow 3Mn_{(s)} + 2Al_2O_{3(s)}$

90. ANS:
a) The reaction is endothermic. The enthalpy of the products is higher than the enthalpy of the reactants. This means that energy is absorbed as the reaction proceeds.
b) A c) C d) B

91. ANS:



92. ANS:
Students might include a diagram to show that breaking a solid into smaller pieces provides a greater surface area. Students might also show, on the same diagram, that more particles of the solvent (or other reactants) collide more frequently with the increased surface area. The increase in surface area results in an increase in the number of collisions. According to collision theory, this should result in an increase in reaction rate.

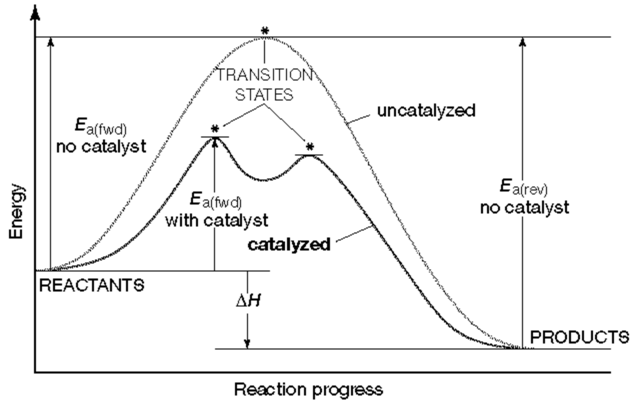


93. ANS:
As the Maxwell-Boltzmann distribution diagram shows, the number of particles with energy that is equal to or greater than the activation energy is increased. This increases the number of successful collisions, thereby increasing the reaction rate.

94. ANS:

As shown in the diagram, an increase in the number of particles of one or more of the reactants increases the probability of collision. The increase in the number of collisions per unit of time results in an increase in reaction rate.

95. ANS:



A catalyst provides an alternate reaction pathway, or reaction mechanism, that has a lower activation energy barrier. Lowering the activation energy barrier results in an increase in the number of particles with sufficient energy to have a successful collision. This results in an increase in the reaction rate.

96. ANS:

Stress	Equilibrium shift
increase in temperature	to the right
increase in hydrogen chloride	to the left
increase in pressure at constant volume	to the left
decrease in volume	to the left
addition of inert gas at constant pressure	to the right

97. ANS:

Moles $C_6H_8O_6$	= Mass ÷ Molar mass
	= 15 g ÷ 176.08 g/mol
	= 0.085 mol
Initial concentration of $C_6H_8O_6$	= Moles ÷ Volume
	= 0.085 mol/1.00 L
	= 0.085 mol/L

Concentration (mol/L)	$C_6H_8O_{6(aq)}$	+	$H_2O_{(l)}$	⇔	$H_3O^+_{(aq)}$	+	$C_6H_7O_6{}^-{}_{(aq)}$
Initial	0.085				0.0		0.0
Change	$-x$				x		x
Final	$0.085 - x$				x		x

$$\frac{[H_3O^+][C_6H_7O_6^-]}{[C_6H_8O_6]}$$

K_a =

$$\frac{x^2}{0.085 - x}$$

=

= 8.0×10^{-5}

Since x is very small, use an approximation.

$$\frac{x^2}{0.085} = 8.0 \times 10^{-5}$$

$x^2 = 6.8 \times 10^{-6}$

$x = 2.6 \times 10^{-3}$

The equilibrium concentration of $H_3O^+_{(aq)}$ is 2.6×10^{-3} mol/L.

pH = $-\log[H_3O^+]$

= $-\log(2.6 \times 10^{-3})$

= 2.59

The pH of the solution is 2.59.

98. ANS:

Mass of 1.0 L of solution	= $D \times V$
	= 1.073 g/mL × 1000 mL
	= 1073 g
Mass HCl	= 10% of 1073 g
	= 107.3 g
Moles HCl	= Mass ÷ Molar mass
	= 107.3 g ÷ 36.46 g/mol
	= 2.94 mol

$[HCl] = [H_3O^+]$

= 2.94 mol/1.0 L

= 2.94 mol/L

pH = $-\log[H_3O^+]$

= $-\log(2.94)$

= -0.47

The pH of the solution is –0.47. (The negative pH indicates that this is a very strong acid.)

99. ANS:

a) Moles Cu = 0.26g ÷ 63.55 g/mol

= 4.09×10^{-3} mol

b) $Cu_{(s)} + 2AgBrO_{3(aq)} \rightarrow 2Ag_{(s)} + Cu(BrO_3)_{2(s)}$

c) Moles $Ag^+_{(aq)}$ = 2 × Moles Cu

= 2 × (4.09 × 10^{-3} m

= 8.18×10^{-3} mol

Concentration of $Ag^+_{(aq)}$ = $\dfrac{8.18 \times 10^{-3}\ \text{mol}}{1.0\text{L}}$

= 8.18×10^{-3} mol/L

d) $K_{sp} = [Ag^+_{(aq)}][BrO_3{}^-_{(aq)}]$

$[BrO_3{}^-_{(aq)}] = [Ag^+_{(aq)}] = 8.18 \times 10^{-3}$ mol/L

$K_{sp} = (8.18 \times 10^{-3}\ \text{mol/L})(8.18 \times 10^{-3}\ \text{mol/L})$

= 6.69×10^{-5}

100. ANS:

a) $Au_{(s)} + NO_3{}^-_{(aq)} + 4H^+_{(aq)} \rightarrow Au^{3+}_{(aq)} + NO_{(g)} + 2H_2O_{(l)}$

0 +5 –2 +1 +3 +2 –2 +1 –2

b) The reactant oxidized is $Au_{(s)}$. The reactant reduced is $NO_3{}^-_{(aq)}$.