SAMPLE AP CHEMISTRY EXAM QUESTIONS

The following questions have been taken from past AP Chemistry Exams and grouped into sections that closely correspond to the chapters in the Fourth Edition of Chemistry by Zumdahl. In some cases questions have been edited from their original format.

Chapter 3 Stoichiometry

1. A sample of dolomitic limestone containing only $CaCO_3$ and $MgCO_3$ was analyzed.

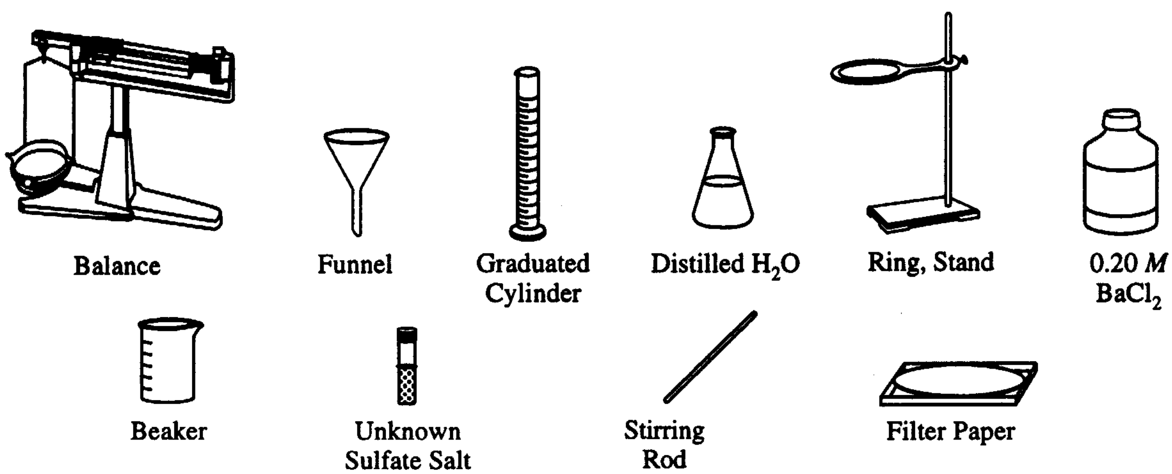
   (a) When a 0.2800 gram sample of this limestone was decomposed by heating, 75.0 mL of CO_2 at 750 mm Hg and 20°C were evolved. How many grams of CO_2 were produced?

   (b) Write equations for the decomposition of both carbonates described above.

   (c) It was also determined that the initial sample contained 0.0448 g of calcium. What percent of the limestone by mass was $CaCO_3$?

   (d) How many grams of the magnesium-containing product were present in the sample in (a) after it had been heated?

2. An experiment is to be performed to determine the mass percent of sulfate in an unknown soluble sulfate salt. The equipment shown below is available for the experiment. A drying oven is also available.



   (a) Briefly list the steps needed to carry out this experiment.

   (b) What experimental data need to be collected to calculate the mass percent of sulfate in the unknown?

   (c) List the calculations necessary to determine the mass percent of sulfate in the unknown.

   (d) Would 0.20 M $MgCI_2$ be an acceptable substitute for the $BaCI_2$ solution provided for this experiment? Explain.

3. Answer the following questions about BeC_2O_4(s) and its hydrate.

(a) Calculate the mass percent of carbon in the hydrated form of the solid that has the formula $BeC_2O_4{\cdot}3H_2O$

(b) When heated to 220.°C, $BeC_2O_4{\cdot}3H_2O$(s) dehydrates completely as represented below,

$$BeC_2O_4{\cdot}3H_2O(s) \rightarrow BeC_2O_4\,(s) + 3\,H_2O\,(g)$$

If 3.21 g of $BeC_2O_4{\cdot}3H_2O$(s) is heated to 220.°C, calculate
(i) the mass of BeC_2O_4 (s) formed, and,
(ii) the volume of the H_2O (g) released, measured at 220.°C and 735 mm Hg

(c) A 0.345 g sample of anhydrous BeC_2O_4 (1) which contains an inert impurity, was dissolved in sufficient water to produce 100. mL of solution. A 20.0 mL portion of the solution was titrated with $KMnO_4$(aq). The balanced equation for the reaction that occurred is as follows.

$$16\,H^+(aq) + 2\,MnO_4^-(aq) + 5\,C_2O_4^{2-(}aq) \rightarrow 2\,Mn^{2+}(aq) + 10\,CO_2(g) + 8\,H_2O(l)$$

The volume of 0.0 150 M KMno4(aq) required to reach the equivalence point was 17.80 mL.

(i) Identify the reducing agent in the titration reaction.

(ii) For the titration at the equivalence point, calculate the number of moles of each of the following that reacted.
- MnO_4^-(aq)
- $C_2O_4^{2-}$(aq)

(iii) Calculate the total number of moles of $C_2O_4^{2-}$(aq) that were present in the 100. mL of prepared solution.

(iv) Calculate the mass percent of BeC_2O_4 (s) in the impure 0.345 g sample.

<u>Chapter 4 Types of Chemical Reactions and Solution Stoichiometry</u>

1. Give the formulas to show the reactants and the products for FIVE of the following chemical reactions. Each of the reactions occurs in aqueous solution unless otherwise indicated. Represent substances in solution as ions if the substance is extensively ionized. Omit formulas for any ions or molecules that are unchanged by the reaction. In all cases a reaction occurs.
<u>Example:</u> A strip of magnesium is added to a solution of silver nitrate.
$$Mg + 2\,Ag^+ \rightarrow Mg^{2+} + 2\,Ag$$
a) Excess sodium cyanide solution is added to a solution of silver nitrate.
b) Solutions of manganese(II) sulfate and ammonium sulfide are mixed.
c) Phosphorus(V) oxide powder is sprinkled over distilled water.
d) Solid ammonium carbonate is heated.
e) Carbon dioxide (,as is bubbled through a concentrated solution of potassium hydroxide.
f) A concentrated solution of hydrochloric acid is added to solid potassium permanganate.
g) A small piece of sodium metal is added to distilled water.
h) A solution of potassium dichromate is added to an acidified solution of iron(II) chloride.

i) Ethanol is burned in oxygen.
j) Solid barium oxide is added to distilled water.
k) Chlorine gas is bubbled into a cold, dilute solution of potassium hydroxide.
l) A solution of iron(II) nitrate is exposed to air for an extended period of time.
m) Excess concentrated sulfuric acid is added to solid calcium phosphate.
n) Hydrogen sulfide gas is bubbled into a solution of mercury(II) chloride.
o) Solid calcium hydride is added to distilled water.
p) A bar of zinc metal is immersed in a solution of copper(II) sulfate.
q) Solutions of tin(II) chloride and iron(III) chloride are mixed.
r) Solutions of cobalt(II) nitrate and sodium hydroxide are mixed.
s) Ethene gas is burned in air.
t) Equal volumes of equimolar solutions of phosphoric acid and potassium hydroxide are mixed.
u) Solid calcium sulfite is heated in a vacuum.
v) Excess hydrochloric acid is added to a solution of diamminesilver(I) nitrate.
w) Solid sodium oxide is added to distilled water.
x) A strip of zinc is added to a solution of 6.0-molar hydrobromic acid.
y) Calcium oxide powder is added to distilled water.
z) Solid ammonium nitrate is heated to temperatures above 300°C.
aa) Liquid bromine is shaken with a 0.5 M sodium iodide solution.
bb) Solid lead(II) carbonate is added to a 0.5 M sulfuric acid solution.
cc) A mixture of powdered iron(III) oxide and powdered aluminum metal is heated strongly.
dd) Methylamine gas is bubbled into distilled water.
ee) Carbon dioxide gas is passed over hot, solid sodium oxide.
ff) A 0.2 M barium nitrate solution is added to an alkaline 0.2 M potassium chromate solution.
gg) A small piece of calcium metal is added to hot distilled water.
hh) Butanol is burned in air
ii) Excess concentrated ammonia solution is added to a solution of nickel(II) sulfate.
jj) A solution of copper(II) chloride is added to a solution of sodium sulfide.
kk) A solution of tin(II) nitrate is added to a solution of silver nitrate.
ll) Excess hydrobromic acid solution is added to a solution of potassium hydrogen carbonate.
mm) Powdered strontium oxide is added to distilled water.
nn) Carbon monoxide gas is passed over hot iron(III) oxide.

4. What will be observed on the surfaces of zinc and silver strips shortly after they are placed in separate solutions of $CuS0_4$**,** as shown below? Account for these observations.

5. Use appropriate chemical principles to explain the observations made when pouring vinegar onto the white residue inside a kettle used for boiling water results in a fizzing/bubbling phenomenon.

6. Concentrated sulfuric acid (18.4 M H_2SO_4) has a density of 1.84 g/mL. After dilution with water to 5.20 M, the solution has a density of 1.38 g/mL and can be used as an electrolyte in lead storage batteries for automobiles.

(a) Calculate the volume of concentrated acid required to prepare 1.00 liter of 5.20 M H_2SO_4.

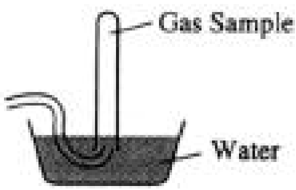
(b) Determine the mass percent of H_2SO_4 in the original concentrated solution.

(c) Calculate the volume of 5.20 M H_2SO_4 that can be completely neutralized with 10.5 grams of sodium bicarbonate, $NaHCO_3$.

(d) What is the <u>molality</u> of the 5.20 M H_2SO_4 ?

Chapter 5 Gases

1. A student collected a sample of hydrogen gas by the displacement of water as shown by the diagram below.



The relevant data are given in the following table.

GAS SAMPLE DATA	
Volume of sample	90.0 mL
Temperature	25°C
Atmospheric Pressure	745 mm Hg
Equilibrium Vapor Pressure of H_2O (25°C)	23.8 mm Hg

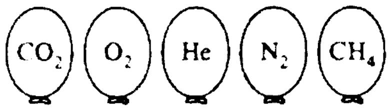
(a) Calculate the number of moles of hydrogen gas collected.

(b) Calculate the number of molecules of water vapor in the sample of gas.

(c) Calculate the ratio of the average speed of the hydrogen molecules to the average speed of the water vapor molecules in the sample.

(d) Which of the two gases, H_2 or H_2O, deviates more from ideal behavior? Explain your answer.

2. Represented below are five identical balloons, each filled to the same volume at 25°C and 1.0 atmosphere pressure with the pure gases indicated.



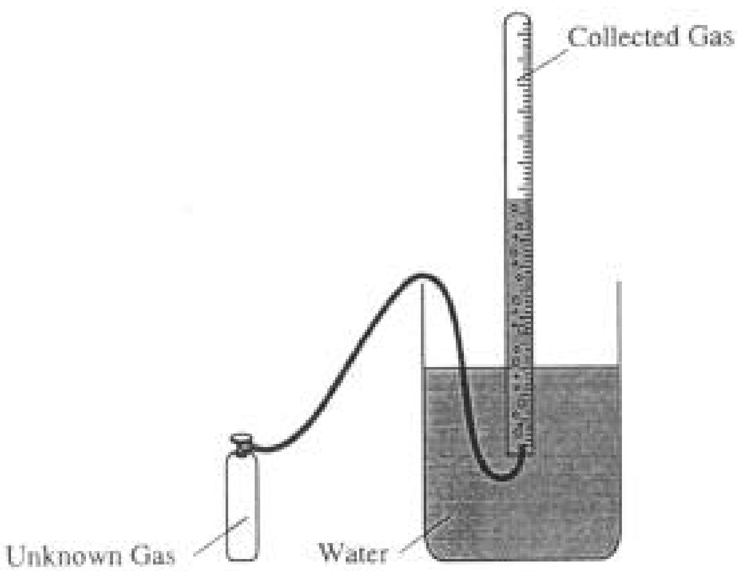
(a) Which balloon contains the greatest mass of gas? Explain.

(b) Compare the average kinetic energies of the gas molecules in the balloons. Explain.

(c) Which balloon contains the gas that would be expected to deviate most from the behavior of an ideal gas? Explain.

(d) Twelve hours after being filled, all the balloons have decreased in size. Predict which balloon will be the smallest. Explain your reasoning.

3. A student performs an experiment to determine the molar mass of an unknown gas. A small amount of the pure gas is released from a pressurized container and collected in a graduated tube over water at room temperature, as shown in the diagram below.



The collection tube containing the gas is allowed to stand for several minutes, and its depth is adjusted until the water levels inside and outside the tube are the same. Assume that:

(i) the gas is not appreciably soluble in water

(ii) the gas collected in the graduated tube and the water are in thermal equilibrium

(iii) a barometer, a thermometer, an analytical balance, and a table of the equilibrium vapor pressure of water at various temperatures are also available.

(a) Write the equation(s) needed to calculate the molar mass of the gas.

(b) List the measurements that must be made in order to calculate the molar mass of the gas.

(c) Explain the purpose of equalizing the water levels inside and outside the gas collection tube.

(d) The student determines the molar mass of the gas to be 64 g mol^{-1}. Write the expression (set-up) for calculating the percent error in the experimental value, assuming that the unknown gas is butane (molar mass 58 g mol^{-1}). Calculations are not required.

(e) If the student fails to use information from the table of the equilibrium vapor pressures of water in the calculation, the calculated value for the molar mass of the unknown gas will be smaller than the actual value. Explain.

Chapter 6 Thermochemistry

1. Propane, C_3H_8, is a hydrocarbon that is commonly used as fuel for cooking.

(a) Write a balanced equation for the complete combustion of propane gas, which yields $CO_2(g)$ and $H_2O(l)$.

(b) Calculate the volume of air at 30°C and 1.00 atmosphere that is needed to burn completely 10.0 grams of propane. Assume that air is 21.0 percent O_2 by volume.

(c) The heat of combustion of propane is -2,220.1 kJ/mol. Calculate the heat of formation, $\Delta H_f°$ of propane given that $\Delta H_f°$ of $H_2O(l)$ = -285.3 kJ/mol and $\Delta H_f°$ of $CO_2(g)$ = -393.5 kJ/mol.

(d) Assuming that all of the heat evolved in burning 30.0 grams of propane is transferred to 8.00 kilograms of water (specific heat = 4.18 J/g˙K), calculate the increase in temperature of the water.

2. When a 2.000 g sample of pure phenol, C_6H_5OH (s), is completely burned according to the equation below, 64.98 kilojoules of heat is released.

$$C_6H_5OH(s) + 7\ O_2\ (g) \rightarrow 6\ CO_2\ (g) + 3\ H_2O\ (l)$$

Use the information in the table below to answer the questions that follow.

Substance	Standard Heat of Formation, $\Delta H°_f$ at 25°C (kJ/mol)	Absolute Entropy, S°, at 25°C (J/mol˙ K)
C(graphite)	0.00	5.69
$CO_2(g)$	-393.5	213.6
$H_2(g)$	0.00	130.6
$H_2O(l)$	-285.85	69.91
O_2 (g)	0.00	205.0
C_6H_5OH (s)	?	144.0

(a) Calculate the molar heat of combustion of phenol in kilojoules per mole at 25°C.

(b) Calculate the standard heat of formation, $\Delta H°_f$, of phenol in kilojoules per mole at 25°C.

(c) If the volume of the combustion container is 10.0 liters, calculate the final pressure in the container when the temperature is changed to 110°C. (Assume no oxygen remains unreacted and that all products are gaseous.)

Chapter 7 Atomic Structure and Periodicity

1. Answer the following questions regarding light and its interactions with molecules, atoms, and ions.

(a) The longest wavelength of light with enough energy to break the Cl-CI bond in $Cl_2(g)$ is 495 nm.

(i) Calculate the frequency, in s^{-1}, of the light.

(ii) Calculate the energy, in J, of a photon of the light.

(iii) Calculate the minimum energy, in kJ mol^{-1}, of the Cl-CI bond.

(b) A certain line in the spectrum of atomic hydrogen is associated with the electronic transition in the H atom from the sixth energy level (n = 6) to the second energy level (n = 2).

(i) Indicate whether the H atom emits energy or whether it absorbs energy during the transition. Justify your answer.

(ii) Calculate the wavelength, in nm, of the radiation associated with the spectral line.

(iii) Account for the observation that the amount of energy associated with the same electronic transition (n = 6 to n = 2) in the He^+ ion is greater than that associated with the corresponding transition in the H atom.

2. Use principles of atomic structure and/or chemical bonding to answer each of the following.

(a) The radius of the Ca atom is 0.197 nanometer; the radius of the Ca^{2+} ion is 0.099 nanometer. Account for this difference.

(b) The lattice energy of CaO (s) is -3,460 kilojoules per mole; the lattice energy for K_20 (s) is -2,240 kilojoules per mole. Account for this difference.

	Ionization Energy (kJ/mol)	
	First	Second
K	419	3,050
Ca	590	1,140

(c) Explain the difference between Ca and K in regard to

(i) their first ionization energies,

(ii) their second ionization energies.

(d) The first ionization energy of Mg is 738 kilojoules per mole and that of Al is 578 kilojoules per mole. Account for this difference.

Chapter 8 Bonding: General Concepts

1. Explain the following in terms of the electronic structure and bonding of the compounds considered.

   (a) Liquid oxygen is attracted to a strong magnet, whereas liquid nitrogen is not.

   (b) The SO_2 molecule has a dipole moment, whereas the C O_2 molecule has no dipole moment. Include the Lewis (electron-dot) structures in your explanation.

   (c) Halides of cobalt(II) are colored, whereas halides of zinc(II) are colorless.

   (d) A crystal of high purity silicon is a poor conductor of electricity; however, the conductivity increases when a small amount of arsenic is incorporated (doped) into the crystal.

2. Explain each of the following observations in terms of the electronic structure and/or bonding of the compounds involved.

   (a) At ordinary conditions, HF (normal boiling point = 20°C) is a liquid, whereas HCI (normal boiling point = -114°C) is a gas.

   (b) Molecules of AsF_3 are polar, whereas molecules of AsF_5 are nonpolar.

   (c) The N-O bonds in the NO_2^- ion are equal in length, whereas they are unequal in HNO_2^-.

   (d) For sulfur, the fluorides SF_2, SF_4, and SF_6 are known to exist, whereas for oxygen only OF_2 is known to exist.

3. Consider the molecules PF_3 and PF_5.

(a) Draw the Lewis electron-dot structures for PF_3 and PF_5 and predict the molecular geometry of each.

(b) Is the PF_3 molecule polar, or is it nonpolar? Explain.

(c) On the basis of bonding principles, predict whether each of the following compounds exists. In each case, explain your prediction.

(i) NF_5 (ii) AsF_5

4. Explain each of the following observations using principles of atomic structure and/or bonding.

(a) Potassium has a lower first ionization energy than lithium.

(b) The ionic radius of N^{3-} is larger than that of O^{2-}.

(c) A calcium atom is larger than a zinc atom.

(d) Boron has a lower first ionization energy than beryllium.

5. Answer the following questions using principles of chemical bonding and molecular structure.

(a) Consider the carbon dioxide molecule, CO_2, and the carbonate ion, CO_3^{2-}.

(i) Draw the complete Lewis electron-dot structure for each species.

(ii) Account for the fact that the carbon-oxygen bond length in CO_3^{2-} is greater than the carbon-oxygen bond length in CO_2 .

(b) Consider the molecules CF_4 and SF_4 .

(i) Draw the complete Lewis electron-dot structure for each molecule.

(ii) In terms of molecular geometry, account for the fact that the CF_4 molecule is nonpolar, whereas the SF_4 molecule is polar.

6. Answer the following questions about the element selenium, Se (atomic number 34).

(a) Samples of natural selenium contain six stable isotopes. In terms of atomic structure, explain what these isotopes have in common, and how they differ.

(b) Write the complete electron configuration (e.g., $ls^2\ 2s^2$... etc.) for a selenium atom in the ground state. Indicate the number of unpaired electrons in the ground-state atom, and explain your reasoning.

(c) In terms of atomic structure, explain why the first ionization energy of selenium is

(i) less than that of bromine (atomic number 35), and

(ii) greater than that of tellurium (atomic number 52).

(d) Selenium reacts with fluorine to form SeF_4. Draw the complete Lewis electron-dot structure for SeF_4 and sketch the molecular structure. Indicate whether the molecule is polar or nonpolar, and justify your answer.

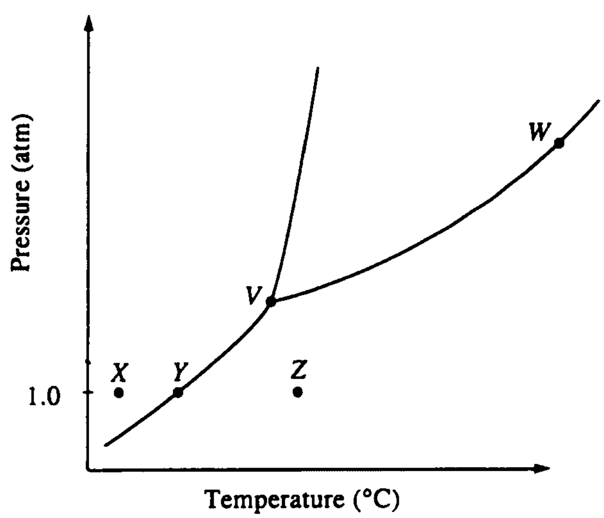
Chapter 10 Liquids and Solids

1. For each of the following, use appropriate chemical principles to explain the observation.

(a) At room temperature, NH_3 is a gas and H_2O is a liquid, even though NH_3 has a molar mass of 17 grams and H_2O has a molar mass of 18 grams.

(b) C (graphite) is used as a lubricant, whereas C (diamond) is used as an abrasive.

2. The phase diagram for a pure substance is shown below.



Use this diagram and your knowledge about changes of phase to answer the following questions.

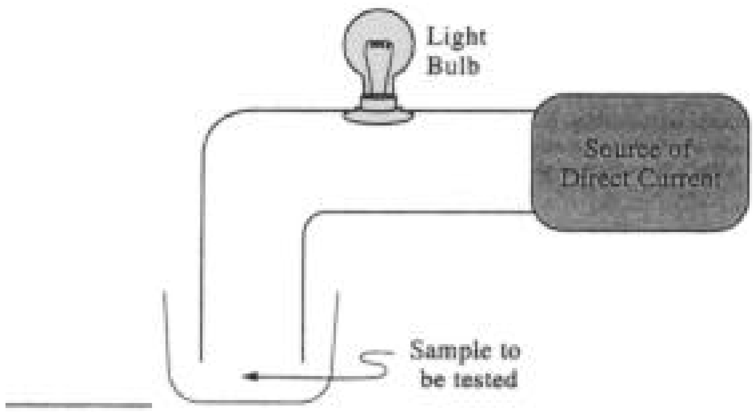
(a) What does point V represent? What characteristics are specific to the system only at point V?

(b) What does each point on the curve between V and W represent?

(c) Describe the changes that the system undergoes as the temperature slowly increases from X to Y to Z at 1.0 atmosphere.

(d) In a solid-liquid mixture of this substance, will the solid float or sink? Explain.

3. The conductivity of several substances was tested using the apparatus represented by the diagram below.



The results of the tests are summarized in the following data table.

	$AgNO_3$	Sucrose	Na	H_2SO_4 (98%)
Melting Point (°C)	2120	1850	990	Liquid at Room Temp.
Liquid (fused)	+ +	-	+ +	+
Water Solution	+ +	-	+ +(1)	+ +(2)
Solid	-	-	+ +	Not Tested

Key: + + Good conductor + Poor conductor - Nonconductor

(1) Dissolves, accompanied by evolution of flammable gas
(2) Conduction increases as the acid is added slowly and carefully to water

Using models of chemical bonding and atomic or molecular structure, account for the differences in conductivity between the two samples in each of the following pairs.

(a) Sucrose solution and silver nitrate solution

(b) Solid silver nitrate and solid sodium metal

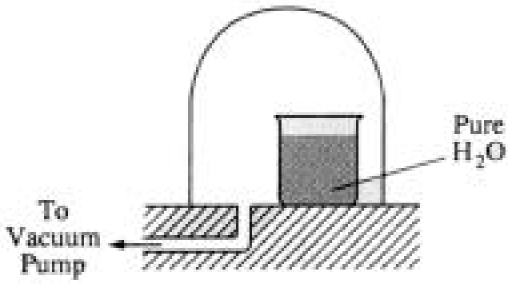
(c) Liquid (fused) sucrose and liquid (fused) silver nitrate

(d) Liquid (concentrated) sulfuric acid and sulfuric acid solution

Chapter 11 Properties of Solutions

1. Discuss the following phenomena in terms of the chemical and physical properties of the substances involved and general principles of chemical and physical change.

   (a) A bell jar connected to a vacuum pump is shown below. As the air pressure under the bell jar decreases, what behavior of water in the beaker will be observed? Explain why this occurs.



   (b) A water solution of I_2 is shaken with an equal volume of a nonpolar solvent such as TTE (trichlorotrifluoroethane). Describe the appearance of this system after shaking. (A diagram may be helpful.) Account for this observation.

2. Use appropriate chemical principles to explain the fact that sodium chloride may be spread on an icy sidewalk in order to melt the ice and that equimolar amounts of calcium chloride are even more effective.

3. An unknown compound contains only the three elements C, H, and O. A pure sample of the compound is analyzed and found to be 65.60 percent C and 9.44 percent H by mass.

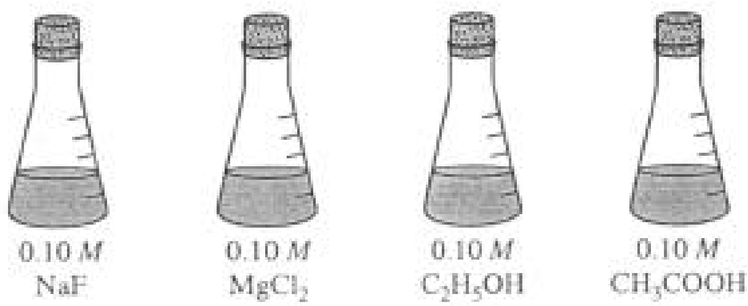
(a) Determine the empirical formula of the compound.

(b) A solution of 1.570 grams of the compound in 16.08 grams of camphor is observed to freeze at a temperature 15.2°C below the normal freezing point of pure camphor. Determine the molar mass and apparent molecular formula of the compound. (The molal freezing-point depression constant, K_f , for camphor is $40.0\ kg\cdot K\cdot mol^{-1}$)

(c) When 1.570 grams of the compound is vaporized at 300°C and 1.00 atmosphere, the gas occupies a volume of 577 milliliters. What is the molar mass of the compound based on this result?

(d) Briefly describe what occurs in solution that accounts for the difference between the results obtained in parts (b) and (c).

4. Answer the following questions, which refer to the 100 rnL samples of aqueous solutions at 25°C in the stoppered flasks shown below.



(a) Which solution has the lowest electrical conductivity? Explain.

(b) Which solution has the lowest freezing point? Explain.

(c) Above which solution is the pressure of water vapor greatest? Explain.

(d) Which solution has the highest pH? Explain.

5. The molar mass of an unknown solid, which is nonvolatile and a nonelectrolyte, is to be determined by the freezing-point depression method. The pure solvent used in the experiment freezes at 10°C and has a known molal freezing-point depression constant, K_f . Assume that the following materials are also available.

• test tubes	• stirrer	• pipet	• thermometer	• balance
• beaker	• stopwatch	• graph paper	• hot-water bath	• ice

(a) Using the two sets of axes provided below, sketch cooling curves for (i) the pure solvent and for (ii) the solution as each is cooled from 20°C to 0.0°C.

(b) Information from these graphs may be used to determine the molar mass of the unknown solid.

(i) Describe the measurements that must be made to determine the molar mass of the unknown solid by this method.

(ii)Show the setup(s) for the calculation(s) that must be performed to determine the molar mass of the unknown solid from the experimental data.

(iii) Explain how the difference(s) between the two graphs in part (a) can be used to obtain information needed to calculate the molar mass of the unknown solid.

(c) Suppose that during the experiment a significant but unknown amount of solvent evaporates from the test tube. What effect would this have on the calculated value of the molar mass of the solid (i.e., too large, too small, or no effect)? Justify your answer.

(d) Show the setup for the calculation of the percentage error in a student's result if the student obtains a value of 126 g·mol^{-1} for the molar mass of the solid when the actual value is 120. g·mol^{-1}.

Chapter 12 Chemical Kinetics

1. Experiments were conducted to study the rate of the reaction represented by the equation below.

$$2\ NO(g) + 2\ H_2(g) \rightarrow N_2(g) \rightarrow 2\ H_20(g)$$

Initial concentrations and rates of reaction are given in the table below.

Experiment	Initial Concentration (mol/L) [NO]	[H_2]	Initial Rate of Formation of N_2 (mol/L.min)
1	0.0060	0.0010	1.8 x 10^{-4}
2	0.0060	0.0020	3.6 x I O^{-4}
3	0.0010	0.0060	0.30 x 10^{-4}
4	0.0020	0.0060	1.2 x 10^{-4}

(a) (i) Determine the order for each of the reactants, NO and H_2, from the data given and show your reasoning.

(ii) Write the overall rate law for the reaction.

(b) Calculate the value of the rate constant, k, for the reaction. Include units.

(c) For experiment 2, calculate the concentration of NO remaining when exactly one-half of the original amount of H_2 had been consumed.

(d) The following sequence of elementary steps is a proposed mechanism for the reaction.

I $\quad NO + NO \Leftrightarrow N_20_2$

II $\quad N_20_2 + H_2 \rightarrow H_20 + N_20$

III $\quad N_20 + H_2 \rightarrow N_2 + H_20$

Based on the data presented, which of the above is the rate-determining step? Show that the mechanism is consistent with

(i) the observed rate law for the reaction, and

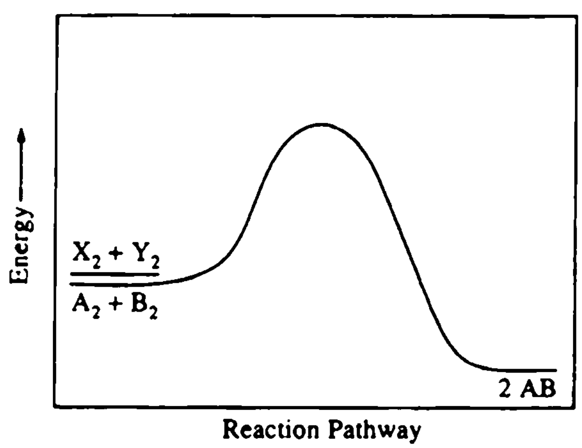
(ii) the overall stoichiometry of the reaction.

2. Two reactions are represented below.

(I) $\quad A_2 + B_2 \rightarrow 2\ AB$

(II) $\quad X_2 + Y_2 \rightarrow 2\ XY$

The potential-energy diagram for reaction I is shown below. The potential energy of the reactants in reaction II is also indicated on the diagram. Reaction II is endothermic, and the activation energy of reaction I is greater than that of reaction II.



(a) Complete the potential-energy diagram for reaction II on the graph above.

(b) For reaction I, predict how each of the following is affected as the temperature is increased by 20°C. Explain the basis for each prediction.

(i) Rate of reaction (ii) Heat of reaction

(c) For reaction II, the form of the rate law is rate = k $[X_2]^m[Y_2]^n$. Briefly describe an experiment that can be conducted in order to determine the values of m and n in the rate law for the reaction.

(d) From the information given, determine which reaction initially proceeds at the faster rate under the same conditions of concentration and temperature. Justify your answer.

3. The reaction between NO and H_2 is believed to occur in the following three-step process.

$NO + NO \Leftrightarrow N_2O_2$ (fast)

$N_2O_2 + H_2 \Leftrightarrow N_2O + H_2O$ (slow)

$N_2O + H_2 \Leftrightarrow N_2 + H_2O$ (fast)

(a) Write a balanced equation for the overall reaction.

(b) Identify the intermediates in the reaction. Explain your reasoning.

(c) From the mechanism represented above, a student correctly deduces that the rate law for the reaction is rate = k $[NO]^2$ $[H_2]$. The student then concludes that

(1) the reaction is third-order and

(2) the mechanism involves the simultaneous collision of two NO molecules and an H_2 molecule.

Are conclusions (1) and (2) correct? Explain.

(d) Explain why an increase in temperature increases the rate constant, k, given the rate law in (c).

4. The following results were obtained when the reaction represented below was studied at 25°C.

$$2A + B \rightarrow C + D$$

Experiment	Initial [Al	Initial [B]	Initial Rate of Formation of C ($mol \cdot L^{-1} \cdot min^{-1}$)
1	0.25	0.75	4.3×10^{-4}
2	0.75	0.75	1.3×10^{-3}
3	1.50	1.50	5.3×10^{-3}
4	1.75	?	8.0×10^{-3}

(a) Determine the order of the reaction with respect to A and to B. Justify your answer.

(b) Write the rate law for the reaction. Calculate the value of the rate constant, specifying units.

(c) Determine the initial rate of change of [A] in Experiment 3.

(d) Determine the initial value of [B] in Experiment 4.

(e) Identify which of the reaction mechanisms represented below is consistent with the rate law developed in part (b). Justify your choice.

1.	$A + B \rightarrow C + M$	Fast
	$M + A \rightarrow D$	Slow
2.	$B \Leftrightarrow M$	Fast equilibrium
	$M + A \rightarrow C + X$	Slow
	$A + X \rightarrow D$	Fast
3.	$A + B \Leftrightarrow M$	Fast equilibrium
	$M + A \rightarrow C + X$	Slow
	$X \rightarrow D$	Fast

5. Answer the following questions regarding the kinetics of chemical reactions.

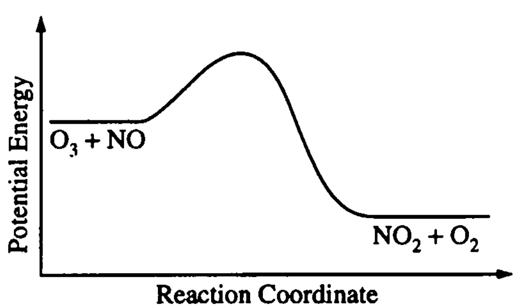
(a) The diagram below shows the energy pathway for the reaction

$$O_3 + NO \rightarrow NO_2 + O_2$$

Clearly label the following directly on the diagram.

(i) The activation energy (Ea) for the forward reaction

(ii) The enthalpy change for the reaction



(b) The reaction

$$2\ N_2O_5 \rightarrow 4\ NO_2 + O_2$$

(i) Sketch the graph that represents the change in $[N_2O_5]$ over time as the reaction proceeds.

(ii) Describe how the graph in (i) could be used to find the reaction rate at a given time, t.

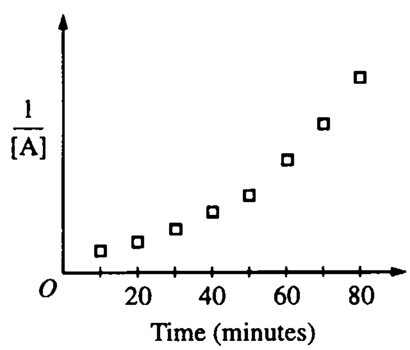
(iii) Considering the rate law and the graph in (i), describe how the value of the rate constant, k, could be determined.

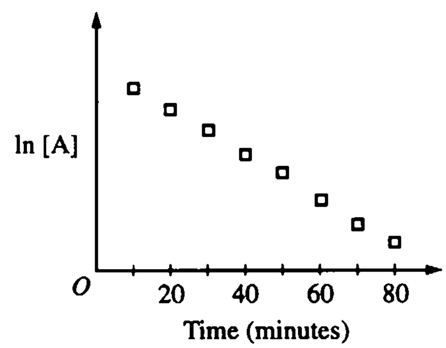
(iv) If more N_2O_5 were added to the reaction mixture at constant temperature, what would be the effect on the Time rate constant, k ? Explain.

(c) Data for the chemical reaction

$$2A \rightarrow B + C$$

were collected by measuring the concentration of A at 10 minute intervals for 80 minutes. The following graphs were generated from analysis of the data.





Use the information in the graphs above to answer the following.

(i) Write the rate-law expression for the reaction. Justify your answer.

(ii) Describe how to determine the value of the rate constant for the reaction.

6. A rate study of the reaction represented below was conducted at 25°C.

$$2NO(g) + Br_2(g) \rightarrow 2NOBr(g)$$

The data that were obtained are shown in the table below.

Experiment	Initial [NO] (mol L^{-1})	Initial [Br_2] (mol L^{-1})	Initial Rate of Appearance of NOBr (mol L^{-1} s^{-1})
1	0.0160	0.0120	3.24×10^{-4}
2	0.0160	0.0240	6.38×10^{-4}
3	0.0320	0.0060	6.42×10^{-4}

(a) Calculate the initial rate of disappearance of $Br_2(g)$ in experiment 1.

(b) Determine the order of the reaction with respect to each reactant, $Br_2(g)$ and NO(g). In each case, explain your reasoning.

(c) For the reaction,

(i) write the rate law that is consistent with the data, and

(ii) calculate the value of the specific rate constant, k, and specify units.

(d) The following mechanism was proposed for the reaction:

$Br_2(g) + NO(g) \rightarrow NOBr_2(g)$ *slow*

$NOBr_2(g) + NO(g) \rightarrow 2NOBr(g)$ *fast*

Is this mechanism consistent with the given experimental observations? Justify your answer.

7. Consider the reaction represented below

$$O_3(g) + NO(g) \rightarrow O_2(g) + NO_2(g)$$

(a) Referring to the data in the table below, calculate the standard enthalpy change, ΔH°, for the reaction at 25°C. Be sure to show your work.

	$O_3(g)$	NO(g)	$NO_2(g)$
Standard enthalpy of formation, $\Delta Hf°$ at 25°C (kJ mol^{-1})	143	90.	33

(b) Make a qualitative prediction about the magnitude of the standard entropy change, ΔS°, for the reaction at 25°C, Justify your answer.

(c) On the basis of your answers to parts (a) and (b), predict the sign of the standard free-energy change,. $\Delta G°$, for the reaction at 25°C. Explain your reasoning.

(d) Use the information in the table below to write the rate-law expression for the reaction, and explain how you obtained your answer.

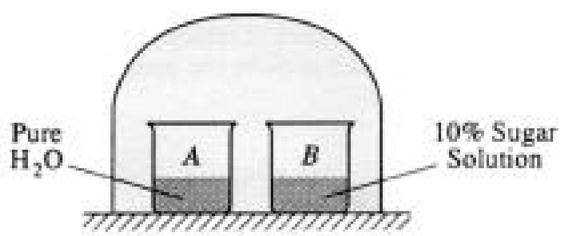
Experiment Number	Initial [O3] ($mol \cdot L^{-1}$)	Initial [NO] ($mol \cdot L^{-1}$)	Initial Rate of Formation of NO2 ($mol \cdot L^{-1} \cdot s^{-1}$)
1	0.0010	0.0010	x
2	0.0010	0.0020	2x
3	0.0020	0.0010	2x
	0.0020	0.0020	4x

(e) The following three-step mechanism is proposed for the reaction. Identify the step that must be the slowest in order for this mechanism to be consistent with the rate-law expression derived in part (d). Explain.

Step 1: $O_3 + NO \rightarrow O + NO_3$
Step 11: $O + O_3 \rightarrow 2\ O_2$
Step III: $NO_3 + NO \rightarrow 2\ NO_2$

Chapter 13 Chemical Equilibrium

1. As the system shown below approaches equilibrium, what change occurs to the volume of water in beaker A? What happens to the concentration of the sugar solution in beaker B? Explain why these changes occur.



2. When H_2 (g) is mixed with $C0_2$ (g) at 2,000 K, equilibrium is achieved according to the equation below.

$$H_2 (g) + CO_2 (g) \Leftrightarrow H_2O (g) + CO (g)$$

In one experiment, the following equilibrium concentrations were measured.

[H] = 0.20 mol/L
$[CO_2 1$ = 0.30 mol/L
$[H_2O1$ = [CO] = 0.55 mol/L

(a) What is the mole fraction of CO (g) in the equilibrium mixture?

(b) Using the equilibrium concentrations given above, calculate the value of *Kc,* the equilibrium constant for the reaction.

(c) Determine Kp in terms of *Kc* for this system.

(d) When the system is cooled from 2,000 K to a lower temperature, 30.0 percent of the CO (g) is converted back to CO_2 (g). Calculate the value of *Kc* at this lower temperature.

(e) In a different experiment, 0.50 mole of H_2 (g) is mixed with 0.50 mole of CO_2 (g) in a 3.0 L reaction vessel at 2,000 K. Calculate the equilibrium concentration, in moles per liter, of CO (g) at this temperature.

3. (a) For the gaseous equilibrium represented below, it is observed that greater amounts of PCI_3 and Cl_2 are produced as the temperature is increased.

$$PCl_5\ (g) \Leftrightarrow PCI_3\ (g) + Cl_2\ (g)$$

(b) If He gas is added to the original reaction mixture at constant volume and temperature, what will happen to the partial pressure of Cl_2 ? Explain.

(c) If the volume of the reaction mixture is decreased at constant temperature to half the original volume, what will happen to the number of moles of Cl_2 in the reaction vessel? Explain.

4. $$C\ (s) + H_2O\ (g) \Leftrightarrow CO\ (g) + H_2\ (g) \qquad \Delta H° = +131\ kJ$$

A rigid container holds a mixture of graphite pellets (C (s)), H_2O (g), CO (g), and H_2 (g) at equilibrium. State whether the number of moles of CO (g) in the container will increase, decrease, or remain the same after each of the following disturbances is applied to the original mixture. For each case, assume that all other variables remain constant except for the given disturbance. Explain each answer with a short statement.

(a) Additional H_2(l) Is added to the equilibrium mixture at constant volume.

(b) The temperature of the equilibrium mixture is increased at constant volume.

(c) The volume of the container is decreased at constant temperature.

(d) The graphite pellets are pulverized.

5. When heated, hydrogen sulfide gas decomposes according to the equation below

$$2\ H_2S(g) \Leftrightarrow 2\ H_2(g) + S_2(g)$$

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 3.72×10^{-2} mol of S_2(g) is present at equilibrium.

(a) Write the expression for the equilibrium constant, *Kc,* 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, *Kc* 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

$$H_2(g) + 1/2\ S_2(g) \Leftrightarrow H_2S(g)$$

at 483 K, calculate the value of the equilibrium constant, *Kc*

Chapter 14 Acids and Bases

1. A chemical reaction occurs when 100 mL of 0.200 M HCI is added drop wise to 100 mL of 0.100 M Na_3PO_4 solution.

(a) Write the two net ionic equations for the formation of the major products.

(b) Identify the species that acts as both a Brönsted acid and as a Brönsted base in the equations in (a). Draw the Lewis electron-dot diagram for this species.

2. Hypochlorous acid, HOCI, is a weak acid commonly used as a bleaching agent. The acid-dissociation constant, Ka for the reaction represented below is 3.2×10^{-8}.

$$HOCI \Leftrightarrow OCI^- + H^+$$

(a) Calculate the $[H^+]$ of a 0.14 M solution of HOCI.

(b) Write the correctly balanced net ionic equation for the reaction that occurs when NaOCI is dissolved in water and calculate the numerical value of the equilibrium constant for the reaction.

(c) Calculate the pH of a solution made by combining 40.0 mL of 0.14 M HOCI and 10.0 mL of 0.56 M NaOH.

(d) How many millimoles of solid NaOH must be added to 50.0 mL of 0.20 M HOCI to obtain a buffer solution that has a pH of 7.49 ? Assume that the addition of the solid NaOH results in a negligible change in volume.

(e) Household bleach is made by dissolving chlorine gas in water, as represented below.

$$Cl_2\,(g) + H_2O \rightarrow H^+ + Cl^- + HOCI\ (aq)$$

Calculate the pH of such a solution if the concentration of HOCI in the solution is 0.065 M.

3. The overall dissociation of oxalic acid, $H_2C_2O_4$, is represented below. The overall dissociation constant is also indicated.

$$H_2C_2O_4 \Leftrightarrow 2\,H^+ + C_2O_4^{2-} \qquad\qquad K = 3.78 \times 10^{-6}$$

(a) What volume of 0.400 M NaOH is required to neutralize completely a 5.00×10^{-3} mole sample of pure oxalic acid?

(b) Give the equations representing the first and second dissociations of oxalic acid. Calculate the value of the first dissociation constant, K_1, for oxalic acid if the value of the second dissociation constant, K_2, is 6.40×10^{-5}.

(c) To a 0.015 M solution of oxalic acid, a strong acid is added until the pH is 0.5. Calculate the $[C_2O_4^{2-}]$ in the resulting solution. (Assume the change in volume is negligible.)

(d) Calculate the value of the equilibrium constant, Kb, for the reaction that occurs when solid $Na_2C_2O_4$ is dissolved in water.

Chapter 15 Applications of Aqueous Equilibria

1. In a saturated solution of MgF_2 at 18°C, the concentration of Mg^{2+} is 1.21×10^{-3} molar. The equilibrium is represented by the equation below.

$$MgF_2\,(s) \Leftrightarrow Mg^{2+}\,(aq) + 2\,F^-\,(aq)$$

(a) Write the expression for the solubility-product constant, *Ksp*, and calculate its value at 18°C.

(b) Calculate the equilibrium concentration of Mg^{2+}, in 1.000 liter of saturated MgF_2 solution at 18°C to which 0.100 mole of solid KF has been added. The KF dissolves completely. Assume the volume change is negligible.

(c) Predict whether a precipitate of MgF_2 , will form when 100.0 milliliters of a 3.00×10^{-3} M $Mg(NO_3)_2$ solution is mixed with 200.0 milliliters of a 2.00×10^{-3} M NaF solution at 18°C. Calculations to support your prediction must be shown.

(d) At 27°C the concentration of Mg^{2+} in a saturated solution of MgF_2 is1. 17 x10^{-3} M. Is the dissolving of MgF_2 in water an endothermic or an exothermic process? Give an explanation to support your conclusion.

2. A chemical reaction occurs when 100 mL of 0.200 M HCI is added drop wise to 100 mL of 0.100 M Na_3PO_4 solution.

(a) Sketch a graph showing the shape of the titration curve that results when 100. mL of the HCI solution is added slowly from a burette to the Na_3PO_4 solution. Account for the shape of the curve.

(b) Write the equation for the reaction that occurs if a few additional milliliters of the HCI solution are added to the solution resulting from the titration in (c).

3. Lead iodide is a dense, golden yellow, slightly soluble solid. At 25°C, lead iodide dissolves in water forming a system represented by the following equation.

$$PbI_2(s) \Leftrightarrow Pb^{2+} + 2\ I^- \quad AH = +46.5 \text{ kilojoules}$$

The solubility-product constant, *Ksp,* for Pbl_2 is 7.1×10^{-9} at 25°C.

(a) If the temperature of the system were lowered from 25°C to 15°C, what would be the effect on the value of *Ksp ?* Explain.

(b) If additional solid PbI_2 were added to the system at equilibrium, what would be the effect on the concentration of $^{I-}$ in the solution? Explain.

4. A 0.500 g sample of a weak, nonvolatile acid, HA, was dissolved in sufficient water to make 50.0 mL of solution. The solution was then titrated with a standard NaOH solution. Predict how the calculated molar mass of HA would be affected (too high, too low, or not affected) by the following laboratory procedures. Explain each of your answers.

(a) After rinsing the burette with distilled water, the burette is filled with the standard NaOH solution; the weak acid HA is titrated to its equivalence point.

(b) Extra water is added to the 0.500 g sample of HA.

(c) An indicator that changes color at pH 5 is used to signal the equivalence point.

(d) An air bubble passes unnoticed through the tip of the burette during the titration.

5. Solve the following problem related to the solubility equilibria of some metal hydroxides in aqueous solution.

(a) The solubility of $Cu(OH)_2$ (s) is 1.72×10^{-6} g/100 mL of solution at 25°C.

(i) Write the balanced chemical equation for the dissociation of $Cu(OH)_2$ (s) in aqueous solution.

(ii) Calculate the solubility (in moles per liter) of $Cu(OH)_2$ at 25°C.

(iii) Calculate the value of the solubility product constant, Ksp, for $Cu(OH)_2$ at 25°C.

(b) The value of the solubility product constant, Ksp, for $Zn(OH)_2$ is 7.7×10^{-17} at 25°C.

(i) Calculate the solubility (in moles per liter) of $Zn(OH)_2$ at 25°C in a solution with a pH of 9.35.

(ii) At 25°C, 50.0 mL of 0.100 M $Zn(NO_3)_2$ is mixed with 50.0 mL of 0.300 M NaOH. Calculate the molar concentration of Zn^{2+} (aq) in the resulting solution once equilibrium has been established. Assume that volumes are additive.

6. An approximately 0.1 M solution of NaOH is to be standardized by titration. Assume that the following materials are available.

- Clean, dry 50 mL burette
- 250 mL Erlenmeyer flask
- Wash bottle filled with distilled water
- Analytical balance
- Phenolphthalein indicator solution
- Potassium hydrogen phthalate (to be used as the primary standard)

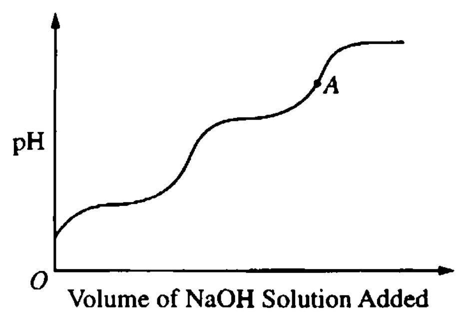
(a) Briefly describe the steps you would take, using the materials listed above, to standardize the NaOH solution.

(b) Describe (i.e., set up) the calculations necessary to determine the concentration of the NaOH solution.

(c) After the NaOH solution has been standardized, it is used to titrate a weak monoprotic acid, HX. The equivalence point is reached when 25.0 mL of NaOH solution has been added. Sketch the titration curve, showing the pH changes that occur as the volume of NaOH solution added increases from 0 to 35.0 mL. Clearly label the equivalence point on the curve.

(d) Describe how the value of the acid-dissociation constant, Ka, for the weak acid HX could be determined from the titration curve in part (c).

(e) The graph below shows the results obtained by titrating a different weak acid, H_2Y, with the standardized NaOH solution. Identify the negative ion that is present in the highest concentration at the point in the titration represented by the letter A on the curve.



7. In aqueous solution, ammonia reacts as represented below.

$$NH_3(aq) + H_20(l) \Leftrightarrow NH_4^+(aq) + OH^-(aq)$$

In 0.0180 M NH_3(aq) at 25°C, the hydroxide ion concentration, [OH^-], is 5.60 x 10^{-4} M. In answering the following, assume that temperature is constant at 25°C and that volumes are additive.

(a) Write the equilibrium-constant expression for the reaction represented above.

(b) Determine the pH of 0.0180 M NH_3(aq).

(c) Determine the value of the base ionization constant, Kb, for NH_3(aq).

(d) Determine the percent ionization of NH_3 in 0.0180 M NH_3(aq).

(e) In an experiment, a 20.0 mL sample of 0.0180 M NH_3(aq) was placed in a flask and titrated to the equivalence point and beyond using 0.0120 M HCl(aq).

(i) Determine the volume of 0.0120 M HCl(aq) that was added to reach the equivalence point.

(ii) Determine the pH of the solution in the flask after a total of 15.0 mL of 0.0120 M HCl(aq) was added.

(iii) Determine the pH of the solution in the flask after a total of 40.0 mL of 0.0120 M HCl(aq) was added.

8. A volume of 30.0 mL of 0.10 *M* NH_3(aq) is titrated with 0.20 *M* HCI(aq). The value of the base-dissociation constant, K_b, for NH_3 in water is 1.8×10^{-5} at 25°C.

(a) Write the net-ionic equation for the reaction of NH_3(aq) with HCI(aq).

(b) Sketch the titration curve that results when a total of 40.0 mL of 0.20 *M* HCI(aq) is added dropwise to the 30.0 mL volume of 0. 10 M NH_3(aq).

(c) From the table below, select the most appropriate indicator for the titration. Justify your choice.

Indicator	**p*Ka***
Methyl Red	5.5
Bromothymol Blue	7.1
Phenolphthalein	8.7

(d) If equal volumes of 0.10 *M* NH_3(aq) and 0.10 *M* NH_4Cl(aq) are mixed, is the resulting solution acidic, neutral, or basic? Explain.

Chapter 16 Spontaneity, Entropy and Free Energy

1. At 298 K, the standard enthalpy change, $\Delta H°$, for the reaction represented below is -145 kilojoules.

$$2\ H_2S\ (g) + SO_2\ (g) \ \Leftrightarrow\ 3\ S\ (s) + 2\ H_2O\ (g)$$

(a) Predict the sign of the standard entropy change, $\Delta S°$, for the reaction. Explain the basis for your prediction.

(b) At 298 K, the forward reaction (i.e., toward the right) is spontaneous. What change, if any, would occur in the value of $\Delta G°$ for this reaction as the temperature is increased? Explain your reasoning using thermodynamic principles.

(c) What change, if any, would occur in the value of the equilibrium constant, Keq, for the situation described in (b)? Explain your reasoning.

(d) The absolute temperature at which the forward reaction becomes non spontaneous can be predicted. Write the equation that is used to make the prediction. Why does this equation predict only an approximate value for the temperature?

2. Lead iodide is a dense, golden yellow, slightly soluble solid. At 25°C, lead iodide dissolves in water forming a system represented by the following equation.

$$PbI_2\ (s) \ \Leftrightarrow\ Pb^{2+} + 2\ I^- \qquad AH = +46.5 \text{ kilojoules}$$

The solubility-product constant, *Ksp*, for Pbl_2 is 7.1×10^{-9} at 25°C.

(a) How does the entropy of the system $PbI_2\ (s) + H_20\ (l)$ change as $PbI_2\ (s)$ dissolves in water at 25°C? Explain.

(b) At equilibrium, $\Delta G = 0$. What is the initial effect on the value of ΔG of adding a small amount of $Pb(NO_3)_2$ to the system at equilibrium? Explain.

3. Information about the substances involved in the reaction represented below

$$C_2H_2\ (g) + 2\ H_2\ (g) \ \rightarrow\ 4\ C_2H_6\ (g)$$

is summarized in the following tables.

Substance	S° (J/mol˙ K)	$\Delta H°_f$ (kJ/mol)
$C_2H_2(g)$	200.9	226.7
$H_2(g)$	130.7	0
$C_2H_6(g)$	?	84.7

Bond	Bond Energy (kJ/mol)
C-C	347
C--C	611
C-H	414
H-H	436

(a) If the value of the standard entropy change, $\Delta S°$, for the reaction is -232.7 J/mole˙Kelvin, calculate the standard molar entropy, S°, of C_2H_6 gas.

(b) Calculate the value of the standard free-energy change, $\Delta G°$, for the reaction. What does the sign of $\Delta G°$ indicate about the reaction above?

(c) Calculate the value of the equilibrium constant, K, for the reaction at 298 K.

(d) Calculate the value of the C≡C bond energy in C_2H_2 in kJ/mol.

4. When a 2.000 g sample of pure phenol, C_6H_5OH (s), is completely burned according to the equation below,

$$C_6H_5OH(s) + 7\ O_2\ (g) \rightarrow 6\ CO_2\ (g) + 3\ H_2O\ (l)$$

64.98 kilojoules of heat is released. Use the information in the table below to answer the questions that follow.

Substance	Standard Heat of Formation, $\Delta H°_f$ at 25°C (kJ/mol)	Absolute Entropy, S°, at 25°C (J/mol˙ K)
C(graphite)	0.00	5.69
CO_2(g)	-393.5	213.6
H_2(g)	0.00	130.6
H_2O(l)	-285.85	69.91
O_2 (g)	0.00	205.0
C_6H_5OH (s)	?	144.0

(a) Calculate the molar heat of combustion of phenol in kilojoules per mole at 25°C.

(b) Calculate the standard heat of formation, $\Delta H°_f$, of phenol in kilojoules per mole at 25°C.

(c) Calculate the value of the standard free-energy change, $\Delta G°$, for the combustion of phenol at 25°C.

5. For the gaseous equilibrium represented below, it is observed that greater amounts of PCI_3 and Cl_2 are produced as the temperature is increased.

$$PCl_5\ (g) \Leftrightarrow PCI_3\ (g) + Cl_2\ (g)$$

(a) What is the sign of $\Delta S°$ for the reaction? Explain.

(b) What change, if any, will occur in $\Delta G°$ for the reaction as the temperature is increased? Explain your reasoning in terms of thermodynamic principles.

6. Answer the following questions in terms of thermodynamic principles and concepts of kinetic molecular theory.

(a) Consider the reaction represented below, which is spontaneous at 298 K.

$$CO_2(g) + 2NH_3(g) \rightarrow CO(NH_2)_2(s) + H_2O(1)\ \Delta H_{298}° = -134kJ$$

(i) For the reaction, indicate whether the standard entropy change, ΔS_{298}, is positive, or negative, or zero. Justify your answer.

(ii) Which factor, the change in enthalpy, ΔH_{298}, or the change in entropy, ΔS_{298}, provides the principal driving force for the reaction at 298 K? Explain.

(iii) For the reaction, how is the value of the standard free energy change, $\Delta G°$, affected by an increase in temperature? Explain.

(b) Some reactions that are predicted by their sign of $\Delta G°$ to be spontaneous at room temperature do not proceed at a measurable rate at room temperature.

(i) Account for this apparent contradiction.

(ii) A suitable catalyst increases the rate of such a reaction. What effect does the catalyst have on $\Delta G°$ for the reaction? Explain.

Chapter 17 Electrochemistry

1. $$Sr(s) + Mg^{2+} \Leftrightarrow Sr^{2+} + Mg(s)$$

Consider the reaction represented above that occurs at 25°C. All reactants and products are in their standard states. The value of the equilibrium constant, *Keq*, for the reaction is 4.2×10^{17} at 250°C.

(a) Predict the sign of the standard cell potential, E°, for a cell based on the reaction. Explain your prediction.

(b) Identify the oxidizing agent for the spontaneous reaction.

(c) If the reaction were carried out at 600°C instead of 25°C, how would the cell potential change? Justify your answer.

(d) How would the cell potential change if the reaction were carried out at 250°C with a 1.0 M solution of $Mg(NO_3)_2$ and a 0. I 0 M solution of $Sr(NO_3)_2$? Explain.

(e) When the cell reaction in (d) reaches equilibrium, what is the cell potential?

2. In an electrolytic cell, a current of 0.250 A is passed through a solution of a chloride of iron, producing Fe (s) and Cl_2 (g).

(a) Write the equation for the half-reaction that occurs at the anode.

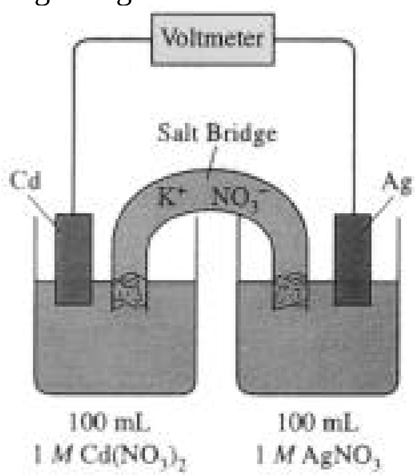
(b) When the cell operates for 2.00 hours, 0.521 g of iron is deposited at one electrode. Determine the formula of the chloride of iron in the original solution.

(c) Write the balanced equation for the overall reaction that occurs in the cell.

(d) How many liters of Cl_2(g), measured at 25°C and 750 mm Hg, are produced when the cell operates as described in part (b) ?

(e) Calculate the current that would produce chlorine gas from the solution at a rate of 3.00 g per hour.

3. Answer the following questions regarding the electrochemical cell shown below.



(a) Write the balanced net-ionic equation for the spontaneous reaction that occurs as the cell operates, and determine the cell voltage.

(b) In which direction do anions flow in the salt bridge as the cell operates? Justify your answer.

(c) If 10.0 mL of 3.0 M $AgNO_3$ solution is added to the half-cell on the right, what will happen to the cell voltage? Explain.

(d)If 1.0 gram of solid NaCl is added to each half-cell, what will happen to the cell voltage? Explain.

(e) If 20.0 mL of distilled water is added to both half-cells, the cell voltage decreases. Explain.

4. Answer each of the following using appropriate chemical principles.

(a) Why does it take longer to cook an egg in boiling water at high altitude than it does at sea level?

(b) When NH_3 gas is bubbled into an aqueous solution of $CuCl_2$, a precipitate forms initially. On further bubbling, the precipitate disappears. Explain these two observations.

(c) Dimethyl ether, $H_3C\text{-}O\text{-}CH_3$, is not very soluble in water. Draw a structural isomer of dimethyl ether that is much more soluble in water and explain the basis of its increased water solubility.

(d) Identify a chemical species that is

(i) capable of oxidizing Cl^- (aq) under standard conditions

(ii) capable of reducing Cl_2 (aq) under standard conditions.

In each case, justify your choice.

5. Answer the following questions that relate to electrochemical reactions.

(a) Under standard conditions at 25°C, Zn(s) reacts with Co^{2+}(aq) to produce Co(s).

(i) Write the balanced equation for the oxidation half reaction.

(ii) Write the balanced net-ionic equation for the overall reaction.

(iii) Calculate the standard potential, E', for the overall reaction at 25°C.

(b) At 25°C, H_20_2 decomposes according to the following equation.

$$2\ H_2O_2(aq) \rightarrow 2\ H_2O(l) + O_2(g) \qquad E° = 0.55\ V$$

(i) Determine the value of the standard free energy change, $\Delta G°$, for the reaction at 25°C.

(ii) Determine the value of the equilibrium constant, *Keq*, for the, reaction at 25°C.

(iii) The standard reduction potential, *E*°, for the half reaction

$$O_2(g) + 4\ H^+(aq) + 4\ e^- \rightarrow 2\ H_2O(1)$$

has a value of 1.23 V. Using this information in addition to the information given above, determine the value of the standard reduction potential, *E*°, for the half reaction below.

$$O_2(g) + 2\ H^+(aq) + 2\ e^- \rightarrow H_2O_2(aq)$$

(c) In an electrolytic cell, Cu(s) is produced by the electrolysis of $CuSO_4$(aq). Calculate the maximum mass of Cu(s) that can be deposited by a direct current of 100. amperes passed through 5.00 L of 2.00 M $CuSO_4$(aq) for a period of 1.00 hour.

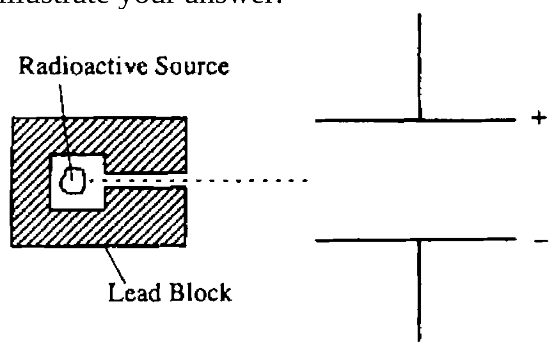
Chapter 21 The Nucleus: A Chemist's Point of View

1. Answer each of the following questions regarding radioactivity.

(a) Write the nuclear equation for decay of $^{239}_{94}$Pu by alpha emission.

(b) Account for the fact that the total mass of the products of the reaction in part (a) is slightly less than that of the original $^{239}_{94}$Pu.

(c) Describe how α, β, and γ rays each behave when they pass through an electric field. Use the diagram below to illustrate your answer.



(d) Why is it not possible to eliminate the hazard of nuclear waste by the process of incineration?