CHEMISTRY PRACTICE TEST 1

Multiple Choice Questions Section I

90 minutes

You may not use a calculator on this section.

Directions: For each of the following questions or incomplete statements, select the best answer or completion from the choices given.

Questions 1–3 refer to the following data table that lists the solubilities, in grams of solute per 100 grams of H_2O, of various salts at two different temperatures.

	Salt	20°C	60°C
I	$Ce_2(SO_4)_3 \cdot 9H_2O$	9.16	3.73
II	KNO_3	31.6	110.0
III	NaCl	36.0	37.3
IV	$K_2Cr_2O_7$	13.1	50.5

1. For which of these salts will the solution process have a ΔH closest to zero?
   A) I
   B) II
   C) III
   D) IV

2. Which of these salts dissolve(s) exothermically?
   A) I only
   B) II only
   C) II, III, and IV only
   D) III and IV only

3. For which of these salts is entropy, but not enthalpy, the driving force for dissolution?
   A) I
   B) II only
   C) III only
   D) II, III, and IV

4. Complete combustion of a compound containing only carbon, hydrogen, and oxygen yields data that allow for elemental analysis. The analysis of the combustion products relies on which of these assumptions?
   I. The quantity of carbon dioxide formed is related directly to the amount of carbon present in the sample.
   II. The quantity of water formed is related directly to the amount of hydrogen present in the sample.

III. The quantity of water formed is limited by the amount of oxygen present in the sample.
IV. The quantity of carbon dioxide formed is limited by the amount of air present.

A) I only
B) II only
C) I, II, III, and IV
D) I and II only

5. Atom Y has 3 valence electrons and atom Z has 6 valence electrons. What is the simplest formula expected for the binary ionic compound composed of Y and Z?

A) Y_2Z
B) YZ_2
C) Y_2Z_3
D) Y_3Z_2

6. The table shows various bond dissociation energies in kJ/mol to two significant figures. Estimate the molar heat of formation of gaseous ammonia.

H—H	N—H	N≡N
450	400	950

A) −50 kJ/mol
B) −100 kJ/mol
C) −550 kJ/mol
D) −1100 kJ/mol

7. One liter of oxygen gas, O_2, and 3 L of sulfur dioxide gas, SO_2, react to form gaseous sulfur trioxide, SO_3, at a given temperature and pressure. How many liters of $SO_3(g)$ can be produced at the same temperature and pressure?

A) 1
B) 2
C) 3
D) 4

8. Considering effective nuclear charge, predict which element will have the smallest atomic radius.

A) H
B) He
C) Li
D) Be

9. Which molecules are polar?

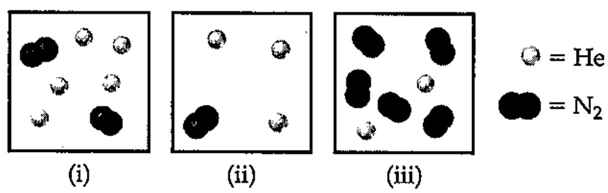
I. H_2O
II. CO_2
III. NO_2
IV. SO_2

A) I and III only
B) I, II, and III only
C) I, II, III, and IV
D) I, III, and IV only

10. Which of the following species is in the greatest concentration in an aqueous 0.100 molar solution of H_3PO_4?

A) H_3O^+
B) H_3PO_4
C) $H_2PO_4^-$
D) HPO_4^{2-}

Questions 11–15 relate to the following information.
Consider the following equal-volume gas samples at the same temperature and under conditions where they all behave as ideal gases.



11. Which gas sample has the highest density?

A) i
B) ii
C) iii
D) All are the same

12. Which gas sample has the highest average kinetic energy of particles?

A) i
B) ii
C) iii
D) All are the same

13. Which gas sample has the highest average velocity of particles?

A) i
B) ii
C) iii
D) All are the same.

14. Which gas sample has the collisions with the highest energy per unit time with the walls of the container?

A) ii only
B) iii only
C) i and iii only
D) All are the same.

15. Assume that conditions change such that all the gas samples behave as real gases. Which gas sample deviates the most from ideal behavior?
    A) i
    B) ii
    C) iii
    D) All are the same.

16. The electronegativity values are shown for two of four elements.

O ---	F 4.0
S 2.5	Cl ---

Select the true statement about the relative values of electronegativity for oxygen and chlorine.
    A) O > Cl because coulombic forces are higher for the smaller O.
    B) O > Cl because the effective nuclear charge is higher for O.
    C) Cl > O because the effective nuclear charge is higher for Cl.
    D) Cl and O are about equal because difference in effective nuclear charge roughly offsets size difference.

17. For 0.1 *M* solutions:

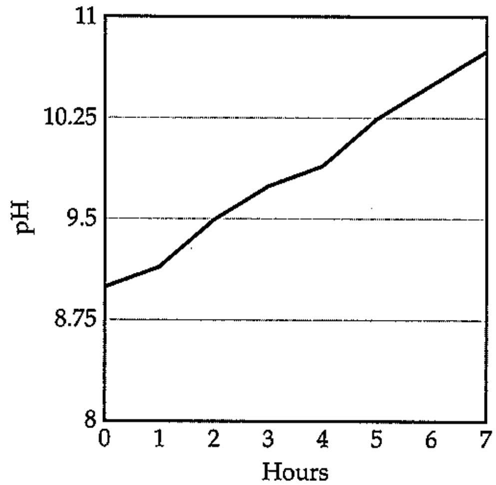
$KHSO_3$	$NaHCO_3$
pH = 5	pH = 9

Which of the following explains these observations?
    A) Potassium ions are more acidic than sodium ions.
    B) Sulfur is more electronegative than carbon.
    C) The HSO_3^- ion is a better proton acceptor than the HCO_3^- ion.
    D) Carbon is smaller than sulfur.

18. A drop of HCl is added to a drop of sodium hydrogen carbonate, and the mixture effervesces. A drop of NaOH is added to a drop of ammonium nitrate. No bubbles are evident, but a piece of filter paper, wetted with phenolphthalein and held above the drop, turns bright pink. All the aqueous reagent solutions are 1.0 *M*.

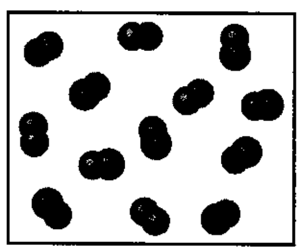
All the following statements are true. What statement best explains why bubbles form in the HCl experiment but not in the NaOH experiment.
    A) Ammonia is polar and CO_2 is nonpolar.
    B) Ammonia is basic and CO_2 is acidic.
    C) CO_2 has a much larger molar mass than ammonia.
    D) Ammonia gas deviates from ideality more than CO_2 gas.

19. The pH of a freshly made 1.0 M solution of sodium hydrogen carbonate is monitored for several hours while constantly stirred. The results are shown in the graph. What best explains the data?



A) The solution slowly absorbed oxygen from the air.
B) The solution slowly absorbed carbon dioxide from the air.
C) The solution slowly lost water through evaporation.
D) Hydrogen carbonate ion slowly reacted to form carbonate ion.

20. The diagram represents a collection of reactant molecules. The dark spheres represent N, and the light spheres represent O.

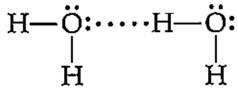


Nitrogen monoxide reacts with oxygen to form nitrogen dioxide. How many nitrogen dioxide molecules would you draw as products if the reaction had a 75% yield?

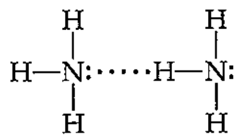
A) 4
B) 6
C) 7
D) 8

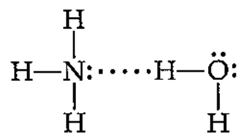
21. Which of the intermolecular forces shown below are the strongest?

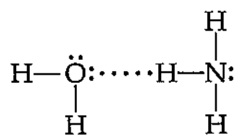
A) v and z only
B) w only
C) x and y only
D) All are equally strong.


(v)

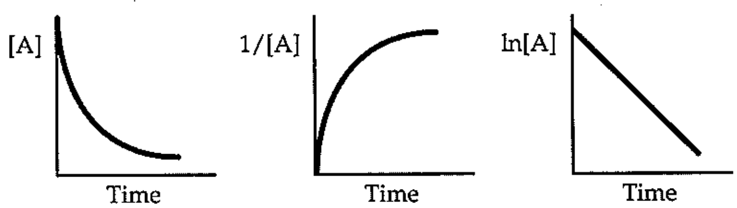

(w)


(x)


(y)


(z)

22. The experimental data from the reaction A → products give these three graphs. What is the most likely order for this reaction?



A) zero
B) first
C) second
D) third

23. The rate of a chemical reaction between substances A and B is found to follow the rate equation $= k[X][Y]^2$. If the concentration of Y is halved, what condition would result in keeping the reaction rate constant, assuming no temperature change?

A) [X] is quadrupled.
B) [X] is doubled.
C) [X] is halved.
D) [X] is tripled.

24. A mixture of 1.00 mole of $H_2(g)$ and 1.00 mole of $I_2(g)$ is placed in a 1.00-L flask at a constant temperature and is allowed to come to equilibrium according to the equation
$H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$.

If the equilibrium constant at this temperature is $K_c = 36.0$, what is the molar concentration of $H_2(g)$ in the equilibrium mixture?

A) 0.500 *M*
B) 1.00 *M*
C) 0.750 *M*
D) 0.250 *M*

25. Which substance in aqueous solution is an electrolyte and is basic?

A) NH_4NO_3
B) CH_3COOH
C) CH_3OH
D) KOH

26. The solubility of which compound is pH dependent?

A) CaF_2
B) KNO_3
C) NaCl
D) CH_3OH

27. A measure of 20.0 mL of 0.10 M solutions of each of the following acids are exactly neutralized with 20.0 mL of 0.10 *M* NaOH. Which of the resulting solutions has the highest pH?

A) $HC_2H_3O_2$ $(K_a = 1.8 \times 10^{-5})$
B) HCN $(K_a = 4.9 \times 10^{-10})$
C) HBrO $(K_a = 2.5 \times 10^{-9})$
D) HIO $(K_a = 2.3 \times 10^{-11})$

28. The Brønsted–Lowry theory of acids and bases predicts that which of the following species would act as an acid?

A) NaH
B) NH_4^+
C) Mg_3N_2
D) NH_2^-

Use the following data for the synthesis of ammonia to answer Questions 29–31.

Temperature (°C)	K_p
300	4.34×10^{-3}
400	1.64×10^{-4}
450	4.51×10^{-5}
500	1.45×10^{-5}
550	5.38×10^{-6}
600	2.25×10^{-6}

Variation in K_p with temperature for $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
$\Delta S°_f$ for ammonia gas $= -198.2$ J/mol-K

29. The reaction for the synthesis of ammonia is
    A) endothermic with an increase in entropy.
    B) endothermic with a decrease in entropy.
    C) exothermic with an increase in entropy.
    D) exothermic with a decrease in entropy.

30. Which conditions would be the best to enhance the amount of ammonia formed?
    A) high temperature and high pressure
    B) low temperature and high pressure
    C) high temperature and low pressure
    D) low temperature and low pressure

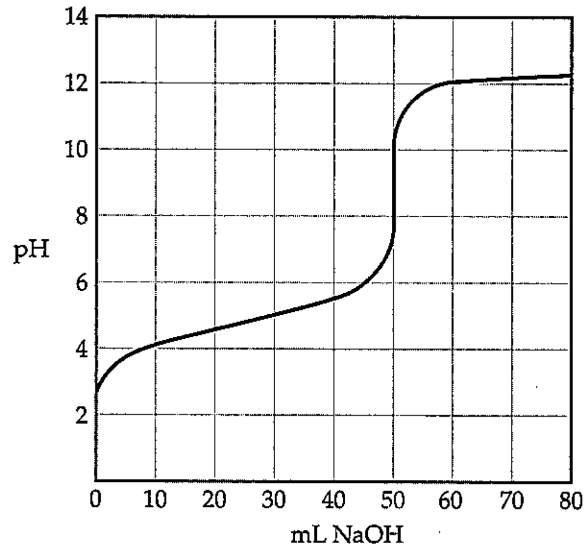
31. Would the use of a catalyst be of practical help in increasing the amount of ammonia formed in the reaction?
    A) No. A catalyst does not change the position of the equilibrium.
    B) No. A catalyst increases the speed of both forward and reverse reactions.
    C) Yes. A catalyst will increase the rate of reaction at lower temperatures.
    D) Yes. A catalyst will increase the rate of reaction and produce more products at all temperatures.

32. Find the value of the equilibrium constant for the following reaction when the equilibrium concentrations are $[N_2] = 2.0\,M$, $[H_2] = 2.0\,M$, $[NH_3] = 2.0\,M$.
    $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
    A) 0.25
    B) 0.50
    C) 0.75
    D) 1.00

33. What is the molar solubility of silver chloride ($K_{sp} = 1.8 \times 10^{-10}$) in 0.050 M sodium chloride?
    A) $(1.8/0.050) \times 10^{-10}\,M$
    B) $(1.8/0.050)^{1/2} \times 10^{-5}\,M$
    C) $(0.050/1.8)^{1/2} \times 10^{-5}\,M$
    D) $(0.050/1.8) \times 10^{-10}\,M$

34. What result is to be expected when 100 mL of a solution that consists of 0.0020 M $Ca(NO_3)_2$ and 0.0020 M $Pb(NO_3)_2$ is mixed with 100 mL of 0.002 M Na_2SO_4? The K_{sp} of $CaSO_4$ is 2.4×10^{-5} and the K_{sp} of $PbSO_4$ is 6.3×10^{-7}.
    A) Both $CaSO_4$ and $PbSO_4$ will precipitate.
    B) Only $CaSO_4$ will precipitate.
    C) Only $PbSO_4$ will precipitate.
    D) Neither $CaSO_4$ nor $PbSO_4$ will precipitate.

Use the figure showing the titration curve for 20.0 mL of analyte titrated with 0.2 *M* titrant to answer Questions 35–40.



35. The figure shows data for the titration of which analyte/titrant pair?
    A) weak acid/weak base
    B) weak acid/strong base
    C) weak base/strong acid
    D) weak base/weak acid

36. Estimate the molar concentration of analyte.
    A) 0.1 *M*
    B) 0.2 *M*
    C) 0.25 *M*
    D) 0.5 *M*

37. Estimate the approximate K_a or K_b for the analyte.
    A) 10^{-3}
    B) 10^{-5}
    C) 10^{-9}
    D) 10^{-12}

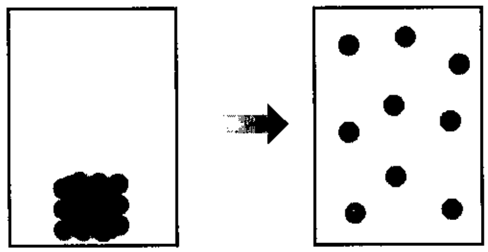
38. Estimate the approximate pH of the equivalence point.
    A) 3
    B) 5
    C) 9
    D) 12

39. Estimate the pH range at which the solution acts as a buffer.
    A) 3–4
    B) 4–6
    C) 11–12
    D) 10–40

40. Estimate the approximate K_a required of a colored indicator to detect a meaningful endpoint of the titration.

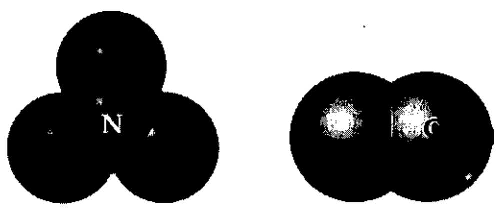
A) 10^{-3}
B) 10^{-5}
C) 10^{-9}
D) 10^{-12}

41. If the process illustrated in the figure happens spontaneously, indicate the signs for ΔH, ΔS, and ΔG, respectively.

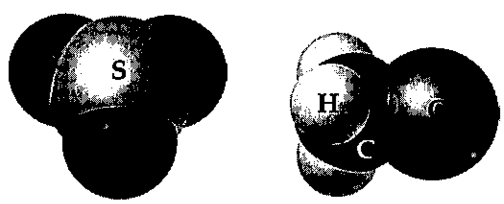


A) $\Delta H = +, \Delta S = +, \Delta G = -$
B) $\Delta H = -, \Delta S = -, \Delta G$ depends on temperature
C) $\Delta H = -, \Delta S = +, \Delta G = -$
D) $\Delta H = +\Delta S = -, \Delta G$ depends on temperature

42. All of the structures pictured below are stable. Which of them could carry a charge and still have an octet Lewis structure?



(i) (ii)



(iii) (iv)

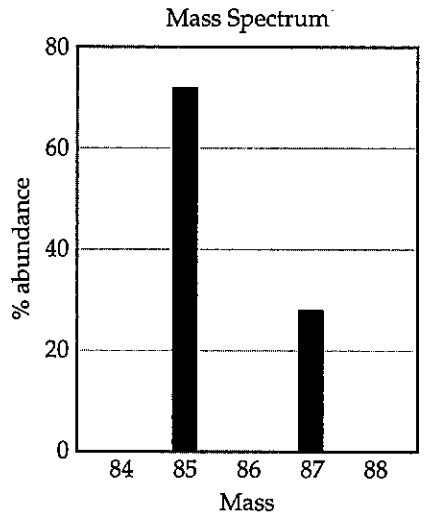
A) i only
B) ii only
C) i and iii only
D) all four

43. Which element can be expected to have the most positive enthalpy for loss of an electron in the gas phase?

A) Al
B) Si
C) P
D) S

44. Which element can be expected to have the most negative enthalpy upon attracting an electron in the gas phase?
    A) B
    B) C
    C) N
    D) O

45. The mass spectrum shown is most likely which of the following?



    A) Rb only
    B) Sr only
    C) a mixture of Rb and Sr
    D) a mixture of At and Fr

46. In the isoelectronic series below, for which species is the least energy required to remove an outer electron?
    A) O^{2-}
    B) F^-
    C) Ne
    D) Na^+

47. Which species below does not violate the octet rule?
    A) BrF_3
    B) BF_3
    C) N_2O
    D) NO_2

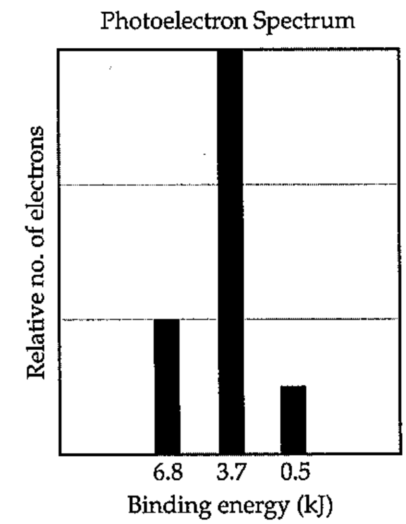
48. How do the properties of the following first-order reaction vary as the reaction goes from reactants to products? $A \rightarrow B$
    A) Rate remains the same and half-life remains the same.
    B) Rate decreases and half-life remains the same.
    C) Rate and half-life both decrease.
    D) Rate decreases and half-life increases.

49. According to the data in this table of initial rates, what is the rate law for the reaction $A + B \rightarrow C + D$

Experiment	[A], *M*	[B], *M*	Rate, *M*/s
1	0.10	0.10	1
2	0.10	0.20	2
3	0.20	0.40	16

A) Rate $= k[A][B]$
B) Rate $= k[A][B]^2$
C) Rate $= k[A]^2[B]$
D) Rate $= k[A]^2[B]^2$

50. This photoelectron spectrum was carried out using 12 kJ photons. The spectrum is most likely to be a partial spectrum of which element?



A) Na
B) Mg
C) Al
D) F

Element	Electron Configuration	Melting Point (°C)	Density	Atomic Radius (Å)	I_1 (kJ/mol)
Fluorine	$[He]2s^22p^5$	–220	1.69 g/L	0.57	1681
Chlorine	$[Ne]3s^23p^5$	–102		1.02	1251
Bromine	$[Ar]4s^23d^{10}4p^5$		3.12 g/cm^3	1.20	1140
Iodine	$[Kr]5s^24d^{10}5p^5$	114	4.94 g/cm^3	1.39	1008

The table above shows some properties of halogens, with some data missing. Use the information in the table to answer Questions 51–54.

51. Predict the approximate melting point of bromine.
    A) −50 °C
    B) 0 °C
    C) 50 °C
    D) 100 °C

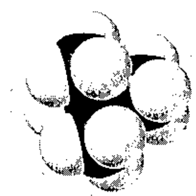
52. Which halogen is most likely to be a liquid at room temperature?
    A) Fluorine
    B) Chlorine
    C) Bromine
    D) Iodine

53. Estimate the approximate density of chlorine at room temperature and pressure.
    A) 1 g/L
    B) 2 g/L
    C) 3 g/L
    D) 4 g/L

54. All the statements below are true. Which statement best explains the trend in the first ionization energies of the halogens?
    A) As effective nuclear charge increases, the first ionization energy increases.
    B) The molar mass decreases from the bottom to the top of a group.
    C) Coulombic attractions increase as valence electrons get closer to the nucleus.
    D) Electronegativity decreases from the top to the bottom of a group.

55. The formulas for *n*-pentane and neopentane are shown below. Which statement is true?

n-Pentane (C_5H_{12})



Neopentane (C_5H_{12})

    A) Neopentane has the higher boiling point because its electrons are more polarizable.
    B) Neopentane has the lower boiling point because it has a greater molar mass.
    C) The *n*-pentane molecule has the higher boiling point because its electrons are more polarizable.
    D) Neopentane has the lower boiling point because it has a lower surface area.

Questions 56–60 relate to the following information.

A voltaic cell is constructed using the following half-reactions.

Half-reaction	***E°***
$Cr^{3+}(aq) + 3\,e^- \rightarrow Cr(s)$	−0.41 V
$PbSO_4(s) + H^+(aq) + 2e^- \rightarrow Pb(s) + HSO_4^-(aq)$	+1.69 V

56. What is the cell voltage?

 A) 1.28 V
 B) 2.10 V
 C) 2.15 V
 D) 4.61 V

57. How many electrons are transferred in the balanced equation for the cell?

 A) 2
 B) 3
 C) 5
 D) 6

58. As the voltaic cell operates, the oxidized and reduced substances are, respectively,

 A) $Cr(s)$ and $PbSO_4(s)$.
 B) Cr^{3+} and $Cr(s)$.
 C) $H^+(aq)$ and $HSO_4^-(aq)$.
 D) $PbSO_4(s)$ and $Cr^{3+}(aq)$.

59. Which change will increase the voltage of the cell?

 A) Increasing the size of the Cr electrode
 B) Increasing the amount of $PbSO_4(s)$
 C) Decreasing the pH
 D) Increasing the $[Cr^{3+}]$

60. Approximately how many grams of metal will be plated at the anode in 100 s at 10 amps?

 A) 1/4
 B) 1/3
 C) 1
 D) 3/2

CHEMISTRY PRACTICE TEST 1

Section II

105 minutes

You may use a calculator for this section.

Directions: Answer each of the following questions, clearly showing the methods you use and the steps involved in arriving at the answers. Partial credit will be given for work shown, and little or no credit will be given for not showing numerical work, even if the answers are correct.

Question 1

a. The acid ionization constants for the triprotic acid known as phosphoric acid are

$K_{a_1} = 7.5 \times 10^{-3}$, $K_{a_2} = 6.2 \times 10^{-8}$, and $K_{a_3} = 4.2 \times 10^{-13}$.

i. Write three ionic equations, one for each of the successive ionizations of the three protons of phosphoric acid. Clearly indicate which equilibrium constant corresponds to each ionic equation.

ii. Calculate the number of grams of sodium dihydrogen phosphate dihydrate needed to prepare 1.00 L of a 0.250 molar solution.

iii. Write a net ionic equation for the hydrolysis of aqueous sodium phosphate, and calculate the value of the corresponding equilibrium constant for the reaction.

b. The acid ionization constant, K_a, for acetic acid is 1.8×10^{-5}. Exactly 41.00 g of sodium acetate is added to water to make 1.00 L of solution. Calculate:

i. the pH of the solution.

ii. the percent ionization of the acetate in the solution.

c. Calculate the pH of a solution resulting from mixing 200.0 mL of the solution prepared in Part b with 200.0 mL of 0.200 M hydrochloric acid.

d. Calculate the number of milliliters of 1.00 M hydrochloric acid or 1.00 M sodium hydroxide (specify which and justify your answer) that needs to be added to 82.00 g of sodium acetate to obtain a buffer having a pH of 4.44.

Question 2

Mercury(II) chloride reacts with oxalate ion according to the following equation:

$$2HgCl_2(aq) + C_2O_4^{2-}(aq) \rightarrow 2Cl^-(aq) + 2CO_2(g) + Hg_2Cl_2(s)$$

The initial rate of the reaction was determined for several concentrations of the reactants, and the following rate data were obtained for the appearance of chloride ion:

Experiment	$[HgCl_2]$ (mol L^{-1})	$[C_2O_4^{2-}]$ (mol L^{-1})	Rate (mol $L^{-1}s^{-1}$)
1	0.144	0.132	5.63×10^{-5}.
2	0.144	0.396	5.10×10^{-4}
3	0.072	0.396	2.51×10^{-4}
4	0.288	0.132	1.13×10^{-4}

a. Write the rate law for this reaction.
b. Calculate the rate constant and specify its units.
c. What is the reaction rate when the concentration of both reactants is 0.150 *M*?
d. What is the rate of disappearance of oxalate ion when $[C_2O_4^{2-}] = 0.10$ and $[HgCl_2] = 0.20\ M$?
e. Is the overall equation likely to be an elementary step? Explain.
f. Which species is oxidized in the reaction and which is reduced?

Question 3

The empirical and molecular formulas of a hydrocarbon are determined by combustion analysis.

a. Combustion of a 1.214-g sample of a hydrocarbon results in 4.059 g of carbon dioxide and 0.9494 g of water.
   i. How many moles of carbon are contained in the sample?
   ii. How many moles of H are contained in the sample?
   iii. What is the empirical formula of the hydrocarbon?
b. The mass spectrum of the hydrocarbon shows a parent peak at 184 mass units.
   i. Predict the molar mass of the hydrocarbon.
   ii. What is the molecular formula of the hydrocarbon?
   iii. Write and balance a chemical equation for the complete combustion of the hydrocarbon.

Question 4

In an experiment, a hydrocarbon and carbon tetrachloride were each found not to dissolve in water.

a. Draw the Lewis structure and a line-angle representation for carbon tetrachloride, and make a sketch of a space-filling model.
b. What is the molecular geometry of carbon tetrachloride and what are the bond angles in the molecule?
c. Does the molecule have any polar bonds? Explain. Is the molecule polar? Explain.

d. Explain the fact that carbon tetrachloride does not dissolve in water.
e. What principal intermolecular force(s) is(are) acting in the carbon tetrachloride solution of the hydrocarbon?

Question 5

Consider the following equilibrium system:

$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g) \quad \Delta H = +176\ \text{kJ/mol}$$

State the effect on the number of grams of solid ammonium chloride present at equilibrium (increase, decrease, or stay the same) when each of the following changes occurs. In each case, explain your reasoning.

a. The partial pressure of ammonia is increased.
b. The temperature is increased.
c. The products are passed through liquid water.
d. The volume of the container is decreased.

Question 6

The first ionization energy of potassium is $+419\ \text{kJ/mol}$. The electron affinity of chlorine is $-349\ \text{kJ/mol}$. Electron affinity is the energy change that occurs when an electron is added to a gaseous atom.

a. Define first ionization energy, and write a thermochemical equation that represents the first ionization energy of potassium.
b. Write a thermochemical equation that represents the electron affinity of chlorine.
c. Write the sum of the reactions you wrote in Parts a and b, and calculate the heat of the reaction in the gas phase. Indicate whether the reaction is endothermic or exothermic.
d. The reaction between potassium metal and chlorine gas to produce solid potassium chloride is highly exothermic ($\Delta H = -435\ \text{kJ/mol}$).
   i. Write the formation reaction of potassium chloride.
   ii. Besides the ionization of potassium, what forces must be overcome in the formation of potassium chloride?
   iii. Explain why this reaction releases so much energy.
e. Predict whether the formation of calcium chloride would be more or less exothermic than the formation of potassium chloride. Explain the basis of your prediction.

Question 7

A 5.0-g sample of sodium chloride is added to 10.0 mL of water at 20 °C. Upon mixing, the salt dissolves and the temperature of the mixture is measured. The experiment is repeated with ammonium chloride and again with calcium chloride. The data in the table summarize the results:

Experiment	Salt	T_1	T_2
1	NaCl	20 °C	20 °C
2	NH_4Cl	20 °C	5 °C
3	$CaCl_2$	20 °C	35 °C

a. Predict the signs of ΔG, ΔS, and ΔH for each of the three experiments. Explain your reasoning.

b. Consider the dissolution of ammonium chloride in water.

i. Write a thermochemical equation, including the heat term, for the dissolution of ammonium chloride.

ii. Discuss the nature and relative magnitudes of the bonds and intermolecular forces that break, and of those that form, when ammonium chloride dissolves in water.

iii. What is the driving force for the change? Use the equation $\Delta G = \Delta H - T\Delta S$ to explain your reasoning.

c. Is there a temperature at which an equilibrium will be established for the dissolution of ammonium chloride? Justify your answer.

MULTIPLE CHOICE ANSWERS AND EXPLANATIONS FOR PRACTICE TEST 1

1. C. The change in solubility of NaCl with temperature is minimal compared to the others, indicating that solid NaCl dissolves in water with little or no heat change.

2. A. As temperature increases, the solubility of cerium sulfate decreases, indicating that cerium sulfate dissolves exothermically. Heating an exothermic process will cause the reaction to shift toward reactants, in this case decreasing solubility.

3. D. The two driving forces that influence chemical and physical change are a decrease in enthalpy and an increase in entropy. Most salts dissolve endothermically, but the accompanying increase in entropy (randomness of the system) overcomes the increase in enthalpy associated with an endothermic reaction. All salts that dissolve endothermically fit into this category.

4. D. The only products obtained from the complete combustion are carbon dioxide and water. The assumptions are that all the carbon in the sample is converted to carbon dioxide, and all the hydrogen in the sample is converted to water. The amounts of carbon and hydrogen can be determined by measuring the masses of carbon dioxide and water produced.

5. C. Atom Y loses three valence electrons to form a Y^{3+} ion. Atom Z gains two electrons to form a Z^{2-} ion. The two ions combine to form Y_2Z_3.

6. A. The formation reaction for one mole of gaseous ammonia is one-half of this balanced equation:

   $$3H_2(g) + N_2(g) \rightleftharpoons 2NH_3(g)$$

   $$\text{Heat of formation} = +\frac{1}{2}[3(\text{H—H}) + (\text{N}\equiv\text{N}) - 6(\text{N—H})]$$

   $$= \frac{1}{2}[3(450) + (950) - 6(400)] = -50\ \text{kJ/mol}$$

7. B. The balanced equation is $O_2(g) + 2SO_2(g) \rightarrow 2SO_3(g)$. Oxygen is the limiting reactant, so twice as many moles of SO_3 can be formed from the available moles of oxygen. At the same conditions of temperature and pressure, liters are proportional to moles.

   $$x\,\text{L}\,SO_3 = 1\,\text{L}\,O_2(2\,\text{L}\,SO_3/1\,\text{L}\,O_2) = 2\,\text{L}\,SO_3$$

8. B. Effective nuclear charge increases from left to right along any period. Neither electron is screened from the charge of the two protons on He, making helium's effective nuclear charge double that of hydrogen. Consequently, helium is smaller than hydrogen.

9. D. Carbon dioxide has two polar bonds, but its linear geometry causes the dipoles to cancel. All the others have polar bonds and have nonlinear geometry.

10. B. Phosphoric acid is a weak polyprotic acid, which ionizes only slightly. All five of the listed species exist in solution, but H_3PO_4 exists to the greatest extent.

11. C. Nitrogen molecules have a higher molar mass than helium atoms.

12. D. Temperature is a measure of the average kinetic energy of molecules. At the same temperature, all three systems have the same average kinetic energy.

13. B. On average, lighter molecules move faster than heavier ones at the same temperature. In System i, 5/7 (20/28) of the molecules are light. In System ii, 3/4 (21/28) of the molecules are light. In System iii, only 2/7 (8/28) of the molecules are light.

14. D. The volumes and temperatures are the same, so the pressures are equal, which means that, on average, the particles hit the walls of the container with equal energy.

15. C. Nitrogen molecules deviate more due to their greater number of electrons and greater polarizability. Nitrogen molecules will form temporary dipoles more often than helium atoms. Thus nitrogen's attractive forces are greater, which makes them stick together more, forming more massive particles.

16. A. The electronegativity values for O and Cl are 3.5 and 3.0, respectively. Oxygen has an effective nuclear charge of about 6+ because eight protons are screened by two inner-core electrons. Chlorine has an effective nuclear charge of about +7 because 17 protons are screened by 10 inner-core electrons. However, chlorine is larger than O, so an electron pair cannot approach the Cl nucleus as closely as it can approach the O nucleus, making the attractive force stronger for O.

17. B. Sulfur's greater electronegativity draws electrons away from the oxygen's in the sulfite ion, making it more stable and allowing the proton to be more easily lost to water.

18. A. Because it is a gas, ammonia escapes the solution. However, ammonia is highly polar and thus can hydrogen-bond with water and is very soluble in water. Therefore, ammonia does not nucleate into bubbles as does nonpolar carbon dioxide, which does not dissolve readily in water.

19. D. Hydrogen carbonate slowly reacts to form carbonate ion with the release of carbon dioxide according to the following equation:

    $2HCO_3^-(aq) \rightarrow H_2O(l) + CO_2(g) + CO_3^{2-}(aq)$.

    The other three statements are all true, but CO_2 is acidic, so the pH would go down with time, and O_2 is neither basic nor acidic, so it would not affect the pH. Evaporation of water would concentrate the hydrogen carbonate base, but not enough to make the pH change by 2 pH units, which represents a 100-fold increase in base concentration.

20. B. The balanced equation is $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$. The diagram shows eight NO molecules and five O_2 molecules available for reaction. Eight NO molecules require four molecules of O_2 to react, whereas five

O_2 molecules would require ten NO molecules. Therefore, NO is the limiting reactant, so theoretically, eight NO molecules will produce eight NO_2 molecules. At 75% yield, there would be six NO_2 molecules formed.

21. B. Because F is smaller and more electronegative than O or N, it is closer to H and has a higher partial negative charge than O or N. Therefore, the coulombic attraction is greatest for H—F.

22. B. A linear plot of ln[A] versus time is characteristic of a simple first-order process. A simple second-order process would yield a straight line for 1/[A] versus time.

23. A. If just [Y] were halved, then the initial rate would decrease by a factor of $(1/2)^2 = 1/4$. To make the rate remain constant, [X] is quadrupled because $(4)^1 = 4$.

24. D . An "ICE" table would look like this:

	$H_2(g)$ +	$I_2(g)$ =	$2HI(g)$
I	1.0 M	1.0 M	0
C	$-x$	$-x$	$+2x$
E	$1.0 - x$	$1.0 - x$	$2x$

$K_c = 36 = (2x)^2/(1.0 - x)(1.0 - x)$

If we take the square root of both sides of the equation, we obtain

$6 = 2x/(1.0 - x)$ or $6 - 6x = 2x, x = 0.750$ $[H_2] = 1 - x = 0.250$

25. D. KOH is a strong base, which dissociates completely in aqueous solution. HCl is a strong acid, CH_3COOH is a weak acid and a weak electrolyte, NH_4NO_3 is a weak acid and a strong electrolyte, and CH_3OH is a neutral nonelectrolyte.

26. A. Fluoride ion is the conjugate base of the weak acid HF. The solubility of calcium fluoride will increase as pH decreases because at low pH the available protons will react with the basic fluoride ion.

$H^+(aq) + F^-(aq) \rightleftharpoons HF(aq)$

Methanol is completely miscible in water, and all the other choices are ionic compounds containing only neutral cations and anions.

27. D. The weakest acid has the strongest conjugate base ($K_w = K_a \times K_b$). At the equivalence point, all that is left is weak base.

28. B. All are bases except NH_4^+. It is unlikely that the already positively charged ammonium ion will accept a proton.

29. D. K decreases with increasing temperature, so the equilibrium is shifted to the left, indicating an exothermic reaction. A negative value for the entropy change indicates a decrease in entropy.

30. B. Low temperature would drive the equilibrium to the right because it is an exothermic reaction. High pressure would also drive the equilibrium to the right because there are fewer moles of gas on the right.

31. C. At lower temperatures the reaction has a higher K, which would increase the amount of ammonia formed. However, at low temperature the rate of the reaction will decrease, so a catalyst is of practical value because it will increase the rate of reaction, even at low temperature.

32. A. $K_c = [NH_3]^2/[N_2][H_2]^3 = (2.0)^2/(2.0)(2.0)^3 = 0.25$

(Remember that no calculators are allowed on the multiple choice section, so the arithmetic required for complex quantitative problems is relatively simple.)

33. A. $AgCl(s) = Ag^+(aq) + Cl^-(aq)$

$K_{sp} = 1.8 \times 10^{-10} = [Ag^+][Cl^-]$

Let $x =$ molar solubility of AgCl.

At equilibrium, $[Ag^+] = x$, $[Cl^-] = 0.050 + x \approx 0.050\,M$

$1.8 \times 10^{-10} = x\,(0.050)$

$x = (1.8/0.050) \times 10^{-10}$

34. C. The concentrations of the ions are not sufficient to form calcium sulfate but are sufficient to form lead(II) sulfate.

$[Pb^{2+}] = 0.0020\,M\,(100/200) = 0.0010\,M$

$[Ca^{2+}] = 0.0020\,M\,(100/200) = 0.0010\,M$

$[SO_4^{2-}] = 0.002\,M\,(100/200) = 0.0010\,M$

$PbSO_4(s) = Pb^{2+}(aq) + SO_4^{2-}(aq)$

$Q = [Pb^{2+}][SO_4^{2-}] = (0.0010)(0.0010) = 1 \times 10^{-6}$

$K_{sp} < Q$, so $PbSO_4(s)$ will form.

$CaSO_4(s) = Ca^{2+}(aq) + SO_4^{2-}(aq)$

$Q = [Ca^{2+}][SO_4^{2-}] = (0.0010)(0.0010) = 1 \times 10^{-6}$

$K_{sp} > Q$, so $CaSO_4(s)$ will not form.

35. B. The pH rises as titrant is added, indicating that the titrant is a base. The high pH at 12 indicates that the base is strong.

36. D. $(20.0\text{ mL})(M) = (50.0\,M)(0.20\,M)$. $M = 0.5\,M$

37. B. At 25 mL of strong base, the pH of the solution equals the pK_a of the weak acid. This pH is approximately 5, so the pK_a is approximately 10^{-5}.

38. C. The equivalence point is approximately the midpoint of the steepest part of the curve, where the pH changes the most.

39. B. A buffer is a solution mixture of a weak acid and its conjugate weak base, both in significant concentration so that the solution resists a change in pH. At about pH $= 4$, significant quantities of the conjugate pair exist until about pH $= 6$.

40. C. The endpoint of a titration is the drop of titrant that causes the indicator, which is a colored weak acid, to change colors. The ideal titration is where the endpoint closely matches the equivalence point. That happens when the indicator has a pK_a that closely approximates the pH of the equivalence point.

41. A. The illustration shows sublimation, an endothermic process with a positive enthalpy change. Gases are more random than solids, so the entropy change is positive. The process happens spontaneously, which means that it is thermodynamically favorable, so the free-energy change is negative.

42. C. Molecule i represents nitrate ion, NO_3^-. Molecule iii can represent sulfur trioxide, SO_3, or sulfite ion, SO_3^{2-}. The others, dichlorine and chloromethane, are neutral molecules.

43. C. In other words, the question asks which element has the highest first ionization energy. First ionization energy increases from left to right along a period because effective nuclear charge also increases. However, sulfur exhibits electron–electron repulsion in a p^4 configuration, making it have a slightly lower ionization energy than phosphorus. This same dip in ionization energy is also seen in N to O and As to Se.

44. D. In other words, the question asks which element has the highest electron affinity. Generally, left to right along any period, electron affinity increases with increasing effective nuclear charge because of increased coulombic attractions. Note: Most elements have negative electron affinities. That is, they attract electrons exothermically, with the release of energy. However, nitrogen has a positive energy associated with its attracting an electron because of electron–electron repulsion when an electron enters nitrogen's p^3 configuration. The electron affinities of the other Group 5A elements, though negative, are also not as negative as the trend would indicate.

45. A. The atomic mass is the average mass of the natural isotopes of an element. The spectrum clearly shows two isotopes of mass 85 (about 70% abundance) and 87 (about 30% abundance). The average would be closer to 85 than 87. The atomic mass of rubidium is about 85.45, and that of strontium is 87.62.

46. A. Oxide ion has just eight protons in its nucleus, two of which are screened by its two 1*s* electrons. Its effective nuclear charge is the smallest among the species listed.

47. C. A valid octet structure can be written for N_2O.

48. B. As a reaction continues, its rate decreases as the concentrations of reactants decrease. The half-life remains the same for any first-order reaction.

49. C. Experiments 1 and 2 show that when [A] is constant and [B] doubles, the rate doubles, meaning the exponent of [B] is 1. Experiments 2 and 3 show that when both [A] and [B] double, the rate increases by a factor of 8. Doubling [B] accounts for a factor of 2, so doubling [A] accounts for a factor of 4. Therefore, the exponent of [A] is 2.

50. A. The spectrum shows decreasing energy along the x-axis, with relative values of 2, 6, and 1 electrons. This could represent an electron configuration of $s^2p^6s^1$. These subshells correspond to three of the four subshells of the sodium atom: $2s^22p^63s^1$. The $1s^2$ subshell is apparently missing because the binding energy of its electrons is greater than the 12 kJ energy of the incident photons.

51. B. There is approximately 100 degrees between the melting points of each halogen.

52. C. With a melting point of 114 °C, iodine is a solid. The low melting points of fluorine and chlorine mean that their boiling points are low also. A melting point of about 0 °C indicates that bromine might still be a liquid at a temperature 20 °C higher.

53. C. Because both fluorine and chlorine are gases at common temperatures, their densities are proportional to their molar masses: $1.69\ \text{g/L} \times 35.5/19 \sim 3\ \text{g/L}$.

54. C. As the size of the atoms decreases up a group, the valence electrons are closer to the effective nuclear charge, so coulombic attractions increase with decreasing distance from the nucleus.

55. D. The more elongated molecule provides a greater surface area, enhances intermolecular contact, and increases dispersion forces. Neopentane is more round, so it has lower surface area and lower dispersion forces.

56. B. The chromium reaction is reversed because it has the lower voltage.
$E^\circ_{\text{cell}} = E^\circ_{\text{ox}} + E^\circ_{\text{red}} = +0.41 + 1.69 = 2.10\ \text{V}$

57. D. To obtain the balanced equation, multiply by coefficients that make the numbers of electrons in the half-reactions cancel:

$$2[Cr(s) \rightarrow Cr^{3+}(aq) + 3\,e^-]$$
$$3[PbSO_4(s) + H^+(aq) + 2e^- \rightarrow Pb(s) + HSO_4^-(aq)]$$

$$2Cr(s) + 3PbSO_4(s) + 3H^+(aq) \rightarrow 2Cr^{3+}(aq) + 3Pb(s) + 3HSO_4^-(aq)$$

58. A. Cr loses 3 electrons and is oxidized. $PbSO_4$ gains 2 electrons and is reduced.

59. C. Decreasing the pH will increase the $[H^+]$ and increase the voltage.

60. C. Round Faraday's constant to ~100,000 C/mol e^- and round the molar mass of lead to ~200 g/mol:

$$x\ \text{g Pb} = 100\ \text{s}(10\ \text{C/s})(1\ \text{mol e}^-/{\sim}100{,}000\ \text{C})(1\ \text{mol Pb}/2\ \text{mol e}^-)$$
$$({\sim}200\ \text{g}/1\ \text{mol}) = {\sim}1\ \text{g Pb}$$

Free Response Answers for Practice Test 1

Answers to Question 1

a. i. $H_3PO_4(aq) + H_2O(l) \rightleftharpoons H_2PO_4^-(aq) + H_3O^+(aq)$

$K_{a_1} = 7.5 \times 10^{-3}$

$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons HPO_4^{2-}(aq) + H_3O^+(aq)$

$K_{a_2} = 6.2 \times 10^{-8}$

$HPO_4^{2-}(aq) + H_2O(l) \rightleftharpoons PO_4^{3-}(aq) + H_3O^+(aq)$

$K_{a_3} = 4.2 \times 10^{-13}$

ii. x g $NaH_2PO_4 \cdot 2H_2O$ = 1.00 L (0.250 mol/L)(156.0 g/mol) = 39.0 g

iii. $PO_4^{3-} + H_2O \rightleftharpoons HPO_4^{2-} + OH^-$

$K_b = \dfrac{K_w}{K_{a_3}} = 1.0 \times 10^{-14}/4.2 \times 10^{-13} = 0.024$

b. i. $CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$

The initial concentration of acetate ion is

(41.00 g)(1 mol/84.0 g)/(1.00 L) = 0.500 M

Let $y = [OH^-]$.

$K_b = \dfrac{K_w}{K_{a_3}} = \dfrac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10} = \dfrac{y^2}{I} = \dfrac{y^2}{0.500}$

$y = [OH^-] = 1.67 \times 10^{-5}$

$pOH = -\log(1.67 \times 10^{-5}) = 4.78$

$pH = 14 - 4.78 = 9.22$

ii. Percent ionization $= 100 \times [OH^-]/I = 100 \times (1.67 \times 10^{-5})/(0.500) = 0.00334\%$

c. $H^+ + CH_3COO^- \rightleftharpoons CH_3COOH$

mol CH_3COO^- = 0.200 L × 0.500 M = 0.100 mol CH_3COO^-

mol H^+ = 0.200 L × 0.200 M = 0.0400 mol H^+

All 0.0400 mol H^+ will react with 0.0400 mol of CH_3COO^-, leaving

0.100 mol − 0.0400 mol = 0.060 mol CH_3COO^- and 0.040 mol CH_3COOH.

$pH = pK_a + \log[CH_3COO^-]/[CH_3COOH] = 4.74 + \log(0.0600/0.040) =$

$4.74 + 0.18 = 4.92$

d. Hydrochloric acid as a limiting reactant needs to be added because sodium acetate is a base. The object is to convert some of the weak base to weak acid, leaving a solution containing both weak acid and base.

$pH = pK_a + \log[CH_3COO^-]/[CH_3COOH]$

$4.44 = 4.74 + \log[CH_3COO^-]/[CH_3COOH]$

$\log[CH_3COO^-]/[CH_3COOH] = -0.30$

$[CH_3COO^-]/[CH_3COOH] = 0.500$

Two-thirds of the base needs to be converted to acetic acid, leaving one-third of the acetate ion unreacted.

82.00 g of sodium acetate is 1 mol.

0.667 mol of HCl is 667 mL of a 1.00 M solution.

Answers to Question 2

a. rate=$k[HgCl_2]^a[C_2O_4^{2-}]^b$

Experiments 1 and 2 show that when $[C_2O_4^{2-}]$ is tripled, the rate goes up by a factor of 9: $5.10 \times 10^{-4}/5.63 \times 10^{-5} = 9.06$, so the value of the exponent $b = 2$ $(3^2 = 9)$. Experiments 1 and 4 show that when $[HgCl_2]$ is doubled, the rate increases by a factor of 2: $1.13 \times 10^{-4}/5.63 \times 10^{-5} = 2.01$, so the exponent $a = 1$ $(2^1 = 2)$.
rate $= k[HgCl_2]^1[C_2O_4^{2-}]^2$

b. rate $= k[HgCl_2]^1[C_2O_4^{2-}]^2$. Use the values of any experiment to obtain k.
$5.63 \times 10^{-5}\ M/s = k(0.144\ M)(0.132\ M)^2$
$k = 0.0224\ 1/M^2\ s$

c. rate $= k[HgCl_2]^1[C_2O_4^{2-}]^2$
rate $= (0.0224\ 1/M^2\ s)(0.150\ M)(0.150\ M)^2 = 7.56 \times 10^{-5}\ M/s$

d. rate $= k[HgCl_2]^1[C_2O_4^{2-}]^2$
rate $= (0.0224\ 1/M^2\ s)(0.20\ M)(0.10\ M)^2 = 4.5 \times 10^{-5}\ M/s$

From the coefficients of the balanced equation, the rate of disappearance of oxalate is half the rate of appearance of chloride ion: rate $= {}^1/_2(4.5 \times 10^{-5}\ M/s)$ $= 2.3 \times 10^{-5}\ M/s$.

e. The overall equation is not likely to be an elementary step because the rate law is third order. An elementary step would require a three-particle collision, which is very rare. More likely, the reaction proceeds via a multistep mechanism that includes only second-order collisions.

f. Oxalate ion, $C_2O_4^{2-}$, is oxidized, and mercury(II) chloride, $HgCl_2$, is reduced.

Answers to Question 3

The empirical and molecular formulas of a hydrocarbon are determined by combustion analysis.

a. i. x mol C = mol CO_2 = (4.059 g)/(44.0 g/mol) = 0.0923 mol C

ii. x mol H = 2 × mol H_2O = 2 × (0.9494 g)/(18.00 g/mol) = 0.1055 mol H

iii. $C_{0.0923}H_{0.1055} = C_{(0.0923)/(0.0932)}H_{(0.1055)/(0.0923)} = C_1H_{1.143} = C_7H_8$

b. i. Molar mass = 184 g/mol

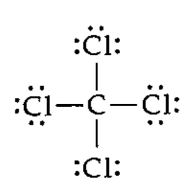
ii. 184.0 g/mol = 2 × 92 g/mol for C_7H_8

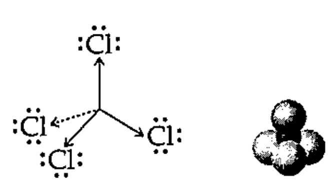
$(C_7H_8) \times 2 = C_{14}H_{16}$

iii. $C_{14}H_{16}(l) + 18O_2(g) \rightarrow 14CO_2(g) + 8H_2O(l)$

Answers to Question 4

a.





b. The molecular geometry is tetrahedral, and the bond angles are all 109.5°.
c. All four C—Cl bonds are polar because of the large difference in electronegativity of carbon and chlorine. The molecule is nonpolar, because the four C—Cl dipoles are oriented in such a way that they cancel each other, giving no net dipole.
d. Water is polar and effectively hydrogen-bonds to other water molecules, excluding any carbon tetrachloride from strong intermolecular attractions to water. Also, the nonpolar carbon tetrachloride molecules have dispersion forces that cause them to exclude the water molecules.
e. The principle forces acting in the carbon tetrachloride–hydrocarbon solution are London dispersions.

Answers to Question 5

$$NH_4Cl(s) \rightleftharpoons NH_3(g) + HCl(g) \Delta H = +176\ \text{kJ/mol}$$

a. The solid will increase because an increased pressure of ammonia will increase the concentration of ammonia, the frequency of collisions, and the rate of the reverse reaction. The equilibrium will rebalance, forming more reactants.
b. The solid will decrease because increasing the temperature favors the endothermic reaction.
c. The solid will decrease because the gaseous products are both soluble in water, which will decrease their concentrations in the gas phase, shifting the equilibrium to the right.
d. The solid will increase because the partial pressures of both gases will increase, causing the reverse reaction to be faster than the forward reaction because of increased collisions.

Answers to Question 6

a. First ionization energy is the energy required to remove the most loosely held electron from the ground state of an isolated gaseous atom.
$419\ \text{kJ} + K(g) \rightarrow K^+(g) + e^-$
b. Electron affinity:
$Cl(g) + e^- \rightarrow Cl^-(g) + 349\ \text{kJ}$
c.
$$\begin{array}{ll} 419\ \text{kJ} + K(g) + \rightarrow K^+(g) + e^- & \Delta H = +419\ \text{kJ/mol} \\ Cl(g) + e^- \rightarrow Cl^-(g) + 349\ \text{kJ} & \Delta H = -349\ \text{kJ/mol} \\ \hline K(g) + Cl(g) \rightarrow K^+(g) + Cl^-(g) & \Delta H = +70\ \text{kJ/mol} \end{array}$$
d. i. $K(s) + {}^1/_2Cl_2(g) \rightarrow KCl(s) \qquad \Delta H = -435\ \text{kJ/mol}$
ii. The Cl — Cl bond must be broken in the formation of potassium chloride. Also, the metallic bonds in solid potassium must be broken.
iii. The attractive force between the positive potassium ion and the negative chloride ion is very strong. The energy released in the formation of the very stable ionic bond is more than enough to compensate for the ionization of potassium and the dissociation of chlorine molecules.
e. The formation of calcium chloride will be more exothermic than the formation of potassium chloride, because the calcium ion has a 2+ charge and the potassium ion has only a 1+ charge. The force of attraction between two oppositely charged ions increases in direct proportion to the magnitude of the charges.

Answers to Question 7

a.

Salt	ΔG	ΔS	ΔH
NaCl	−	+	undetermined
NH_4Cl	−	+	+
$CaCl_2$	−	+	−

All three salts dissolved spontaneously, so ΔG is negative for all three processes.
Dissolution of ionic solids will probably lead to more randomness or disorder, so ΔS is likely to be positive for all three processes.
Because no noticeable temperature change was observed for NaCl, it cannot be determined whether it dissolves endothermically or exothermically.
A decrease in temperature means that ammonium chloride dissolves endothermically.
An increase in temperature means that calcium chloride dissolves exothermically.

b. i. $\text{Heat} + NH_4Cl(s) \rightarrow NH_4^+(aq) + Cl^-(aq)$

ii. Ammonium chloride contains ionic bonds that require heat to break them. Hydrogen bonds between water molecules also break. Upon dissolution, ion–dipole interactions form between water molecules and ammonium and chloride ions, with the release of energy. Because the process is endothermic, more heat is absorbed than released, so the ionic bonds and the broken intermolecular forces of the solvent must be stronger than the ion–dipole interactions.

iii. The driving force for the dissolution of ammonium chloride is the increase in entropy. ΔS for the process has a large positive value.

Considering $\Delta G = \Delta H - T\Delta S$, ΔG is negative because the process happens spontaneously, and ΔH is positive because the system cools off. Thus ΔS must have a relatively large positive value to offset the positive value of ΔH.

c. Yes, the dissolution of NH_4Cl will establish an equilibrium. The equilibrium condition is when $\Delta G = 0$. Consider $\Delta G = \Delta H - T\Delta S$. Both ΔH and ΔS are positive. As the temperature decreases, the $T\Delta S$ term will become less positive until a point is reached when $T\Delta S$ equals ΔH. At this point they cancel, and ΔG equals 0.