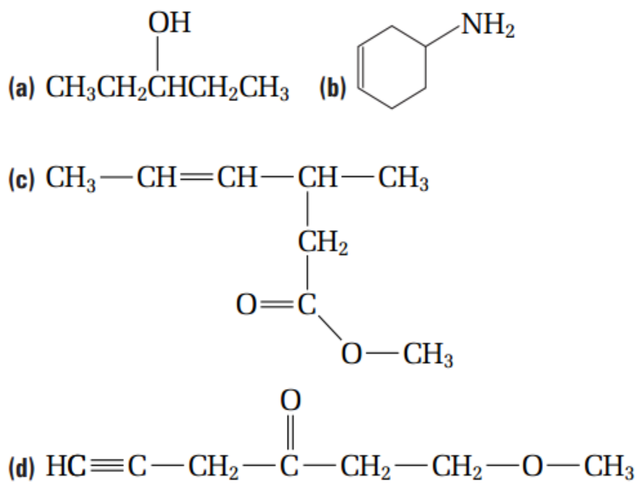
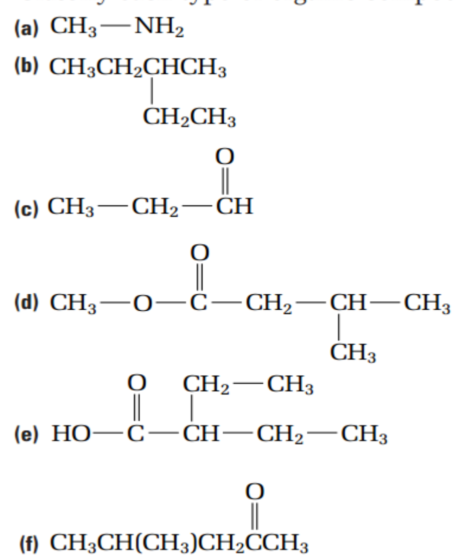
EXAM REVIEW – SCH4U

UNIT 1 - ORGANIC CHEMISTRY

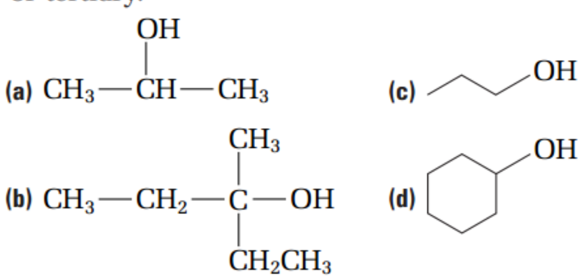
1.
Identify the functional group(s) in each molecule.



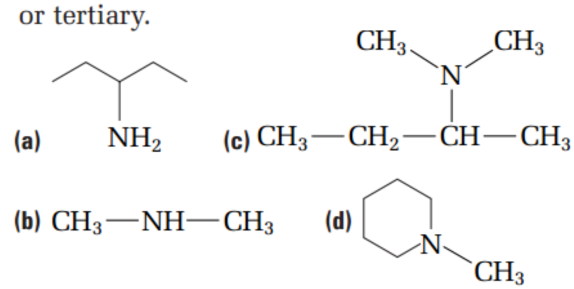
2.
Classify each type of organic compound.



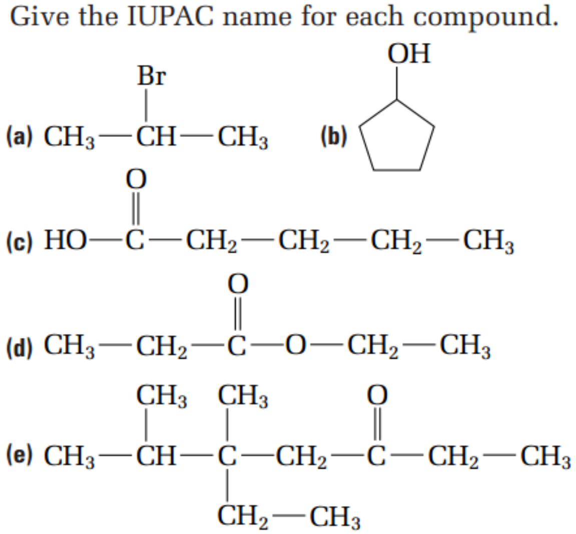
3.
Identify each alcohol as primary, secondary, or tertiary.



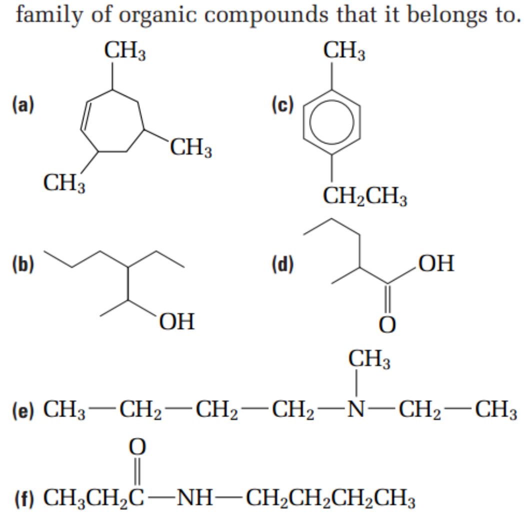
4.
Identify each amine as primary, secondary, or tertiary.



5.
Give the IUPAC name for each compound.



6.
Name each compound. Then identify the family of organic compounds that it belongs to.



7.

Draw a condensed structural diagram for each compound.

(a) 1-propanamine
(b) 3-ethylpentane
(c) 4-heptanol
(d) propanoic acid
(e) cyclobutanol
(f) methoxyethane
(g) 1,1-dibromobutane
(h) 2-methyl-3-octanone
(i) hexanal
(j) N-ethylpropanamide
(k) methyl butanoate

8.

Draw a structural diagram for each compound.

(a) 3,4-dimethylheptanoic acid
(b) 3-ethyl-3-methyl-1-pentyne
(c) N-ethyl-2,2-dimethyl-3-octanamine
(d) N-ethyl-N-methylhexanamide
(e) 1,3-dibromo-5-chlorobenzene

9.

Draw a structural diagram for each compound. Then identify the family of organic compounds that it belongs to.

(a) 4-ethylnonane
(b) 4-propylheptanal
(c) 3,3-dimethyl-2-hexanamine
(d) 2-methoxypentane
(e) para-dimethylbenzene

10.

Identify the error in each name. Then correct the name.

(a) 2-pentanal
(b) 1,3-dimethylpropane
(c) 2,2-dimethylbenzene
(d) N,N-diethyl-N-methylpentanamide
(e) 1-methylpropanoic acid

11.

(a) Draw condensed structural diagrams for five isomers with the molecular formula C_6H_{12}

(b) Draw line structural diagrams for five new isomers that also have the molecular formula C_6H_{12}. Include one pair of cis-trans isomers.

12.

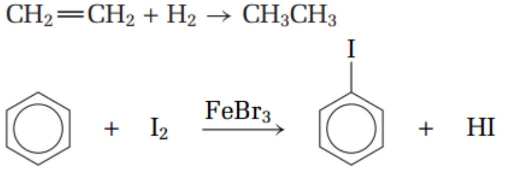
Describe each type of organic reaction, and give an example.

(a) addition
(b) substitution
(c) elimination
(d) oxidation
(e) reduction
(f) condensation
(g) hydrolysis

13.

Identify each reaction as an addition reaction, a substitution reaction, or an elimination reaction.

(a) $CH_2{=}CH_2 + H_2 \rightarrow CH_3CH_3$

(b) 

(c) $CH_3CH_2CH_2OH \rightarrow CH_3CH{=}CH_2 + H_2O$

(d) $CH_3CH_2CH_2OH + HCl \rightarrow CH_3CH_2CH_2Cl + H_2O$

(e) $CH_3CH(CH_3)CH{=}CH_2 + HBr \rightarrow CH_3CH(CH_3)CH(Br)CH_3$

(f) $CH_3CH_2CH{=}CHCH_2CH_3 + H_2O \rightarrow CH_3CH_2CH(OH)CH_2CH_2CH_3$

(g) $CH_3CH_2Br + NH_3 \rightarrow CH_3CH_2NH_2 + HBr$

14.

Describe each type of polymerization, and give an example.

(a) addition polymerization
(b) condensation polymerization

15.

(a) What is Markovnikov's rule? Why does it apply to the following reaction?
$CH_3CH{=}CH_2 + HBr \rightarrow ?$

(b) Name and draw the two isomeric products of this reaction.

(c) Which product is formed in the greater amount?

16.

(a) What is the difference between a protein and an amino acid?

(b) How are proteins important to living organisms?

17.

(a) Define the term "lipid."

(b) Give three different examples of lipids.

(c) List four foods you eat that contain lipids.

(d) How are lipids important to your body?

18.

(a) Distinguish between monosaccharides, disaccharides, and polysaccharides.

(b) Draw and name an example of each.

(c) How are carbohydrates important to living organisms?

19.

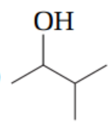
Draw and name the product(s) of each incomplete reaction. (**Hint:** Do not forget to include any second products, such as H_2O or HBr.)

(a) $CH_3CH{=}CHCH_3 + Br_2 \rightarrow$

(b) $HO{-}CH_2CH_2CH_2CH_2CH_3 + HBr \rightarrow$

(c) $CH_3CH_2\overset{\overset{\displaystyle O}{\|}}{C}H + [O] \rightarrow$

(d) $HO{-}CH_2CH_2CH_3 \xrightarrow[\Delta]{H_2SO_4}$

(e)  $+ \ [O] \rightarrow$

(f) $CH_3CH_2\overset{\overset{\displaystyle O}{\|}}{C}OH + HOCH_3 \xrightarrow{H_2SO_4}$

(g) $CH_3CH{=}CHCH_2CH_3 + H_2 \rightarrow$

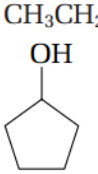
(h) $CH_3{-}\overset{\overset{\displaystyle O}{\|}}{C}{-}O{-}CH_2CH_3 + H_2O \xrightarrow[\Delta]{H_2SO_4}$

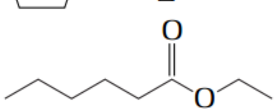
20.

Draw the product(s) of each incomplete reaction. **Hint:** Do not forget to include the second product, such as H_2O or HBr, for a substitution reaction.

(a) $CH_3CH_2{-}\overset{\overset{\displaystyle O}{\|}}{C}{-}OH + CH_3CH_2CH_2OH \rightarrow$

(b) $CH_3CH_2C{\equiv}CH + Cl_2 \rightarrow$ (i) $+ Cl_2 \rightarrow$ (ii)

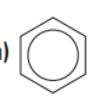
(c)  $\xrightarrow[\Delta]{H_2SO_4}$

(d)  $+ \ H_2O \xrightarrow[\Delta]{H_2SO_4}$

(e) $CH_3{-}\overset{\overset{\displaystyle O}{\|}}{C}{-}\underset{\underset{\displaystyle CH_3}{|}}{CH}{-}CH_3 + [H] \rightarrow$

(f) $H_2C{=}CHCH_2CH(CH_3)_2 + HOH \rightarrow$ (i) (major product) + (ii) (minor product)

(g) $CH_3CH_2CH_2CH_2CH_2{-}OH + [O] \rightarrow$ (i) $+ [O] \rightarrow$ (ii)

(h)  $+ \ Cl_2 \xrightarrow{FeBr_3}$

(i) $CH_3CH_2CH(Br)CH_3 + OH^- \rightarrow$

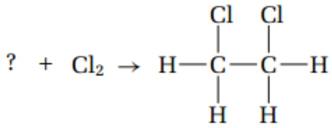
(j) $HO{-}CH_2CH_2CH_2CH_3 + H\overset{\overset{\displaystyle O}{\|}}{C}{-}OH \rightarrow$

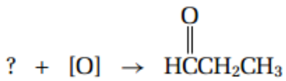
(k) $n CH_2{=}\underset{\underset{\displaystyle OH}{|}}{CH} \xrightarrow{\text{polymerization}}$

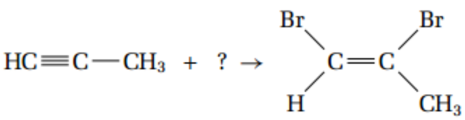
(l) $n HO(CH_2)_7OH + n HO\overset{\overset{\displaystyle O}{\|}}{C}CH_2\overset{\overset{\displaystyle O}{\|}}{C}OH \xrightarrow{\text{polymerization}}$

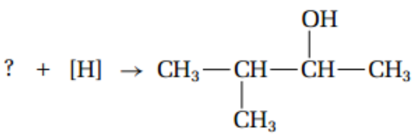
21.

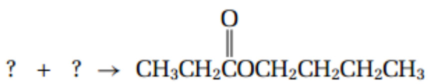
Draw and name the reactant(s) in each reaction.

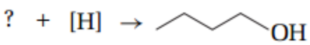
(a) 

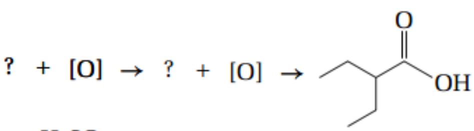
(b) 

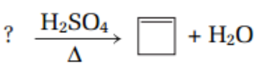
(c) 

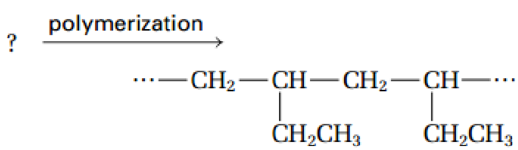
(d) 

(e) 

(f) 

(g) 

(h) 

(i) 

UNIT 2 – ENERGY CHANGES AND RATES OF REACTION

1.

A given chemical equation is tripled and then reversed. What effect, if any, will there be on the enthalpy change of the reaction?

2.

Write the balanced equation for the formation of each substance.

(a) $LiCl_{(s)}$ **(b)** $C_2H_5OH_{(\ell)}$ **(c)** $NH_4NO_{3(s)}$

3.

A 10.0 g sample of pure acetic acid, CH_3CO_2H, is completely burned. The heat released warms 2.00 L of water from 22.3°C to 39.6°C. Assuming that no heat was lost to the calorimeter, what is the enthalpy change of the complete combustion of acetic acid? Express your answer in units of kJ/g and kJ/mol.

4.

Use equations (1), (2), and (3) to find the enthalpy change of the formation of methane, CH_4, from chloroform, $CHCl_3$.

$CHCl_{3(\ell)} + 3HCl_{(g)} \rightarrow CH_{4(g)} + 3Cl_{2(g)}$

(1) $\frac{1}{2}H_{2(g)} + \frac{1}{2}Cl_{2(g)} \rightarrow HCl_{(g)} \quad \Delta H^\circ = -92.3 \text{ kJ}$

(2) $C_{(s)} + 2H_{2(g)} \rightarrow CH_{4(g)} \quad \Delta H^\circ = -74.8 \text{ kJ}$

(3) $C_{(s)} + \frac{1}{2}H_{2(g)} + \frac{3}{2}Cl_{2(g)} \rightarrow CHCl_{3(\ell)} \quad \Delta H^\circ = -134.5 \text{ kJ}$

5.

The following equation represents the combustion of ethylene glycol, $(CH_2OH)_2$.

$(CH_2OH)_{2(\ell)} + \frac{5}{2}O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(\ell)}$

$\Delta H^\circ = -1178$ kJ

Use known enthalpies of formation and the given enthalpy change to determine the enthalpy of formation of ethylene glycol.

6.

The following equation represents the complete combustion of butane, C_4H_{10}.

$C_4H_{10(g)} + 6.5O_{2(g)} \rightarrow 4CO_{2(g)} + 5H_2O_{(g)}$

(a) Using known enthalpies of formation, calculate the enthalpy change of the complete combustion of C_4H_{10}. (The enthalpy of formation of C_4H_{10} is –126 kJ/mol.)

(b) Using known enthalpies of formation, calculate the enthalpy change of the complete combustion of ethane, C_2H_6, to produce carbon dioxide and water vapour. Express your answer in units of kJ/mol and kJ/g.

(c) A 10.0 g sample that is 30% C_2H_6 and 70% C_4H_{10}, by mass, is burned in excess oxygen. How much energy is released?

7.

At elevated temperatures, ammonia reacts with oxygen as follows:

$4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(g)}$

(a) Write an equation that shows the relationship between the rate of reaction expressed in terms of each reactant and product.

(b) The average rate of production of nitrogen monoxide is 6.2×10^{-2} mol/(L • s). What is the average rate of change in the concentration of ammonia?

8.

State two requirements for an effective collision between reactants.

9.

State the difference between a homogeneous catalyst and a heterogeneous catalyst.

10.

Phosgene, $COCl_2$, is a highly toxic gas that is heavier than air. It can be produced by reacting carbon monoxide with chlorine in a very slow reaction.

$CO_{(g)} + Cl_{2(g)} \rightarrow COCl_{2(g)}$

The following initial rate data were collected at a particular temperature.

Experiment	Initial [CO] (mol/L)	Initial [Cl_2] (mol/L)	Initial rate (mol/(L • s))
1	0.500	0.0500	6.45×10^{-30}
2	0.0500	0.0500	6.65×10^{-31}
3	0.0500	0.500	6.50×10^{-30}
4	0.05 00	0.00500	6.60×10^{-32}

(a) Write the rate law equation for this reaction.

(b) Calculate the value of the rate constant. Make sure that you use the proper units.

11.

The following reaction was studied using the method of initial rates.

$3A_{(aq)} + 4B_{(aq)} \rightarrow$ products

The following data were collected.

Experiment	Initial [A] (mol/L)	Initial [B] (mol/L)	Initial rate (mol/(L • s))
1	0.200	0.200	5.00
2	0.600	0.200	45.0
3	0.200	0.400	10.0
4	0.600	0.400	90.0

(a) Write the rate law equation.

(b) What is the overall reaction order?

(c) Calculate the value of the rate constant, with the proper units.

12.

. A chemical reaction between compounds C and D is first order in C and second order in D. Find the unknown information in the table below.

Experiment	Rate (mol/(L • s))	[A] (mol/L)	[B] (mol/L)
1	0.10	1.0	0.20
2	(a)	2.0	0.20
3	(b)	2.0	0.40

13.

Consider the reaction below.

$2A + B \rightarrow C + D$

At 20°C, the activation energy of the forward reaction is 59.9 kJ/mol and the activation energy of the reverse reaction is 72.0 kJ/mol.

(a) What is the enthalpy change for the reaction?

(b) Sketch a potential energy diagram for the reaction.

14.

. Consider the reaction below.

$2A + B_2 \xrightarrow{C} D + E$

A chemist proposes the following reaction mechanism.

Step 1 $A + B_2 \rightarrow AB_2$

Step 2 $AB_2 + C \rightarrow AB_2C$

Step 3 $AB_2C + A \rightarrow A_2B_2 + C$

Step 4 $A_2B_2 \rightarrow D + E$

Suggest a rate law equation corresponding to each of the following situations. Remember that a rate law equation may include only the concentration of reactants and catalyst, if any. If you think that it is impossible to predict a rate law for any of the situations, explain why. State any assumptions you make for each situation.

(a) Step 1 is the rate determining step.

(b) Step 2 is the rate determining step.

(c) Step 3 is the rate determining step.

15.

(a) Explain, in your own words, what is meant by the term "activation energy."

(b) How can the idea of activation energy be used to explain the temperature dependence of rate?

(c) How can activation energy be used to explain why a catalyst increases the rate of a chemical reaction?

UNIT 3 – CHEMICAL SYSTEMS AND EQUILIBRIUM

1.

Name the factors that can affect the equilibrium of a reaction.

2.

The following reaction is at equilibrium. Which condition will produce a shift to the right: a decrease in volume or a decrease in temperature? Explain why.

$H_{2(g)} + Cl_{2(g)} \rightleftharpoons 2HCl_{(g)} + \text{heat}$

3.

The following system is at equilibrium. Will an increase in pressure result in a shift to the left or to the right? How do you know?

$2CO_{2(g)} \rightleftharpoons 2CO + O_{2(g)}$

4.
The oxidation of sulfur dioxide to sulfur trioxide is an important reaction. At 1000 K, the value of K_c is 3.6×10^{-3}.

$2SO_{2(g)} + O_{2(g)} \rightleftharpoons 2SO_{3(g)}$

A closed flask originally contains 1.7 mol/L $SO_{2(g)}$ and 1.7 mol/L $O_{2(g)}$. What is $[SO_3]$ at equilibrium when the reaction vessel is maintained at 1000 K?

5.
0.50 mol of $CO_{(g)}$ and 0.50 mol of $H_2O_{(g)}$ are placed in a 10 L container at 700 K. The following reaction occurs.

$CO_{(g)} + H_2O_{(g)} \rightleftharpoons H_{2(g)} + CO_{2(g)}$ $\quad K_c = 8.3$

What is the concentration of each gas that is present at equilibrium?

6.
The following results were collected for two experiments that involve the reaction, at 600°C, between gaseous sulfur dioxide and oxygen to form gaseous sulfur trioxide. Show that the value of K_c was the same in both experiments.

Experiment 1		Experiment 2	
Initial concentration (mol/L)	Equilibrium concentration (mol/L)	Initial concentration (mol/L)	Equilibrium concentration (mol/L)
$[SO_2] = 2.00$	$[SO_2] = 1.50$	$[SO_2] = 0.500$	$[SO_2] = 0.590$
$[O_2] = 1.50$	$[O_2] = 1.25$	$[O_2] = 0$	$[O_2] = 0.0450$
$[SO_3] = 3.00$	$[SO_3] = 3.50$	$[SO_3] = 0.350$	$[SO_3] = 0.260$

7.
Equal amounts of hydrogen gas and iodine vapour are heated in a sealed flask.

(a) Sketch a graph to show how $[H_{2(g)}]$ and $[HI_{(g)}]$ change over time.

(b) Would you expect a graph of $[I_{2(g)}]$ and $[HI_{(g)}]$ to appear much different from your first graph? Explain why.

8.
Give two examples of each of the following acids and bases.

(a) Arrhenius acids

(b) Brønsted-Lowry bases

(c) Brønsted-Lowry bases that are not Arrhenius bases

9.
Write a chemical formula for each acid or base.

(a) the conjugate base of OH^-

(b) the conjugate acid of ammonia, NH_3

(c) the conjugate acid of HCO_3^-

(d) the conjugate base of HCO_3^-

10.
. In each pair of bases, which is the stronger base?

(a) $HSO_4^-{}_{(aq)}$ or $SO_4^{2-}{}_{(aq)}$

(b) $S^{2-}{}_{(aq)}$ or $HS^-{}_{(aq)}$

(c) $HPO_4^{2-}{}_{(aq)}$ or $H_2PO_4^-{}_{(aq)}$

(d) $HCO_3^-{}_{(aq)}$ or $CO_3^{2-}{}_{(aq)}$

11.
Sodium methanoate, NaHCOO, and methanoic acid, HCOOH, can be used to make a buffer solution. Explain how this combination resists changes in pH when small amounts of acid or base are added.

12.
A student dissolved 5.0 g of vitamin C in 250 mL of water. The molar mass of ascorbic acid is 176 g/mol, and its K_a is 8.0×10^{-5}. Calculate the pH of the solution. **Note:** Abbreviate the formula of ascorbic acid to H_{Asc}.

13.
During an experiment, a student pours 25.0 mL of 1.40 mol/L nitric acid into a beaker that contains 15.0 mL of 2.00 mol/L sodium hydroxide solution. Is the resulting solution acidic or basic? What is the concentration of the ion that causes the solution to be acidic or basic?

14.

If the pH of urine is outside the normal range of values, this can indicate medical problems. Suppose that the pH of a urine sample was measured to be 5.53 at 25°C. Calculate pOH, $[H_3O^+]$, and $[OH^-]$ for the sample.

15.

Propanoic acid, CH_3CH_2COOH, is a weak monoprotic acid that is used to inhibit mould formation in bread. A student prepared a 0.10 mol/L solution of propanoic acid and found that the pH was 2.96. What is the acid dissociation constant for propanoic acid? What percent of its molecules were dissociated in the solution?

16.

Formic acid, HCOOH, is present in the sting of certain ants. What is the pH of a 0.025 mol/L solution of formic acid?

17.

Pyridine, C_5H_5N, is used to manufacture medications and vitamins. Calculate the base dissociation constant for pyridine if a 0.125 mol/L aqueous solution has a pH of 9.10.

18.

Sodium acetate, CH_3COONa, is used for developing photographs. Find the value of K_b for the acetate ion. Then calculate the pH of a solution that contains 12.5 g of sodium acetate dissolved in 1.00 L of water. (Only the acetate ion affects the pH of the solution.)

19.

Calculate the molar solubility of zinc hydroxide at 25°C, where K_{sp} is 7.7 x 10^{-17}.

20.

If 150 mL of a 0.200 mol/L $CaCl_{2(aq)}$ and 150 mL of 0.05 mol/L $Na_2SO_{4(aq)}$ are mixed at 20°C, determine whether a precipitate will form. For $CaSO_{4(aq)}$ at 20°C, K_{sp} is 3.6 x 10^{-5}.

21.

What is the molar solubility of $PbCl_{2(s)}$ in a 0.4 mol/L $NaCl_{(aq)}$ solution at SATP?

22.

In titration of 30.00 mL of 0.400 mol/L $HC_2H_3O_{2(aq)}$ with standardized 0.400 mol/L $NaOH_{(aq)}$, what is the amount of unreacted $HC_2H_3O_{2(aq)}$ and the pH of the solution:

a) before titration begins
b) during titration but before the equivalence point (15.00 mL of 0.400 mol/L $NaOH_{(aq)}$ added)
c) at the equivalence point (30.00 mL of 0.400 mol/L $NaOH_{(aq)}$ added

23.

A 1.0 L buffer is prepared that contains 0.40 mol/L acetic acid and 0.40 mol/L sodium acetate at equilibrium.
a) calculate the pH of the buffer
b) If 0.10 mol of $H^+_{(aq)}$ is added to the buffer without changing its volume, calculate the pH. Assume no volume change.
c) Compare the change in pH to the change expected if the same amount of $H^+_{(aq)}$ is added to water to make a 1.0 L solution.

UNIT 4 - ELECTROCHEMISTRY

1.
Determine the oxidation number of each element present in the following substances.

(a) BaH_2
(b) Al_4C_3
(c) KCN
(d) $LiNO_2$
(e) $(NH_4)_2C_2O_4$
(f) S_8
(g) AsO_3^{3-}
(h) VO_2^+
(i) XeO_3F^-
(j) $S_4O_6^{2-}$

2.
Determine which of the following balanced chemical equations represent redox reactions. For each redox reaction, identify the oxidizing agent and the reducing agent.

(a) $2C_6H_6 + 15O_2 \rightarrow 12CO_2 + 6H_2O$
(b) $CaO + SO_2 \rightarrow CaSO_3$
(c) $H_2 + I_2 \rightarrow 2HI$
(d) $KMnO_4 + 5CuCl + 8HCl \rightarrow KCl + MnCl_2 + 5CuCl_2 + 4H_2O$

3.
Determine which of the following balanced net ionic equations represent redox reactions. For each redox reaction, identify the reactant that undergoes oxidation and the reactant that undergoes reduction.

(a) $2Ag^+_{(aq)} + Cu_{(s)} \rightarrow 2Ag_{(s)} + Cu^{2+}_{(aq)}$
(b) $Pb^{2+}_{(aq)} + S^{2-}_{(aq)} \rightarrow PbS_{(s)}$
(c) $2Mn^{2+} + 5BiO_3^- + 14H^+ \rightarrow 2MnO_4^- + 5Bi^{3+} + 7H_2O$

4.
(a) Examples of molecules and ions composed only of vanadium and oxygen are listed below. In this list, identify molecules and ions in which the oxidation number of vanadium is the same.

V_2O_5
V_2O_3
VO_2
VO
VO_2^+
VO^{2+}
VO_3^-
VO_4^{3-}
$V_3O_9^{3-}$

(b) Is the following reaction a redox reaction?
$2NH_4VO_3 \rightarrow V_2O_5 + 2NH_3 + H_2O$

5.
Use the half-reaction method to balance each of the following equations.

(a) $MnO_2 + Cl^- \rightarrow Mn^{2+} + Cl_2$ (acidic conditions)
(b) $NO + Sn \rightarrow NH_2OH + Sn^{2+}$ (acidic conditions)
(c) $Cd^{2+} + V^{2+} \rightarrow Cd + VO_3^-$ (acidic conditions)
(d) $Cr \rightarrow Cr(OH)_4^- + H_2$ (basic conditions)
(e) $S_2O_3^{2-} + NiO_2 \rightarrow Ni(OH)_2 + SO_3^{2-}$ (basic conditions)
(f) $Sn^{2+} + O_2 \rightarrow Sn^{4+}$ (basic conditions)

6.
Use the oxidation number method to balance each of the following equations.

(a) $SiCl_4 + Al \rightarrow Si + AlCl_3$
(b) $PH_3 + O_2 \rightarrow P_4O_{10} + H_2O$
(c) $I_2O_5 + CO \rightarrow I_2 + CO_2$
(d) $SO_3^{2-} + O_2 \rightarrow SO_4^{2-}$

7.

Calculate the standard cell potential for the galvanic cell in which the following reaction occurs.

$2I^-_{(aq)} + Br_{2(\ell)} \rightarrow I_{2(s)} + 2Br^-_{(aq)}$

8.

Calculate the standard cell potential for the galvanic cell in which the following reaction occurs.

$2Na_{(s)} + 2H_2O_{(\ell)} \rightarrow 2NaOH_{(aq)} + H_{2(g)}$

9.

Predict whether each reaction is spontaneous or non-spontaneous under standard conditions.

(a) $Cd_{(s)} + Cu^{2+}_{(aq)} \rightarrow Cd^{2+}_{(aq)} + Cu_{(s)}$ **(b)** $I_{2(s)} + 2Cl^-_{(aq)} \rightarrow 2I^-_{(aq)} + Cl_{2(g)}$

10. Explain the function of the following parts of an electrolytic cell. **(a)** electrodes **(b)** electrolyte **(c)** external voltage	11. In a galvanic cell, one half-cell has a cadmium electrode in a 1 mol/L solution of cadmium nitrate. The other half-cell has a magnesium electrode in a 1 mol/L solution of magnesium nitrate. Write the shorthand representation.
12. What is the importance of the hydrogen electrode?	13. Write the half-reactions and calculate the standard cell potential for each reaction. Identify each reaction as spontaneous or non-spontaneous. **(a)** $Zn_{(s)} + Fe^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Fe_{(s)}$ **(b)** $Cr_{(s)} + AlCl_{3(aq)} \rightarrow CrCl_{3(aq)} + Al_{(s)}$ **(c)** $2AgNO_{3(aq)} + H_2O_{2(aq)} \rightarrow 2Ag_{(s)} + 2HNO_{3(aq)} + O_{2(g)}$

UNIT 5 – STRUCTURE AND PROPERTIES

1. Explain how the Bohr atomic model differs from the Rutherford atomic model, and explain the observations and inferences that led Bohr to propose his model.	2. Briefly describe the contributions made by the following physicists to the development of the quantum mechanical model of the atom. **(a)** Planck **(b)** de Broglie **(c)** Einstein **(d)** Heisenberg **(e)** Schrödinger
3. Explain how Pauli's exclusion principle and Hund's rule assist you in writing electron configurations.	4. Give the energy level and type of orbital occupied by the electron with the following set of quantum numbers: $n = 3$, $l = 1$, $m_l = 0$, $m_s = +\frac{1}{2}$.

5.

Which of the following is the correct orbital diagram for the third and fourth principal energy levels of a vanadium atom ($Z = 23$)? Justify your answer.

(a) [↑] [↑↓][↑][↑][↑][ ] [ ][ ][ ]

(b) [↑] [↑][↑][↑][↑][↑] [ ][ ][ ]

(c) [↑↓] [↑][↑][↑][ ][ ] [ ][ ][ ]

(d) [↑↓] [↑][↑][↑][↑][ ] [ ][ ][ ]

6.

Each of the following orbital diagrams is incorrect. Identify the errors, explain how you recognized them, and use the aufbau principle to write electron configurations using the corrected orbital diagrams.

(a) carbon: [↑↓] [↑↓] [↑↓][ ][ ]

(b) iron: [↑↓] [↑][↑][↑][↑][↑] [↑][ ][ ]

(c) bromine: [↑↓] [↑↓][↑↓][↑↓][↑↓][↑↓] [↑][↑][↑]

7.

. The electron configurations below represent atoms in excited states. Identify each atom, and write its ground state electron configuration.

(a) $1s^2 2s^2 sp^6 3s^1 3p^1$

(b) $1s^2 2s^2 2p^6 3s^2 3p^4 4s^1$

(c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^4 4p^1$

(d) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1 4d^2$

8.

. Identify elements whose atoms have the following valence electron configurations:

(a) $5s^1$ **(c)** $3s^2$

(b) $4s^2 3d^2$ **(d)** $4s^2 3d^{10} 4p^3$

9.

Why are there no *p* block elements in period 1 of the periodic table?

10.

Describe the intermolecular forces between the molecules of hydrogen halides (HF, HCl, HBr, HI), and explain the difference in their boiling points.

11.

What is the difference between a permanent molecular dipole and an induced dipole in a non-polar molecule?

12.

What types of intermolecular forces must be broken to melt solid samples of the following?

(a) NH_3 **(b)** NaI **(c)** Fe **(d)** CH_4

13.

In which compound, H_2O or in NH_3, will the hydrogen bonding be stronger? Explain.

14.

Determine the molecular shape of the hydronium ion, H_3O^+.

15.

Determine the shape of SiF_6^{2-} using VSEPR theory.

16.

Use VSEPR theory to predict the molecular shape for each of the following:

(a) HCN **(b)** SO_2 **(c)** SO_3 **(d)** SO_4^{2-}

17.

Use VSEPR theory to predict the molecular shape for each of the following:

(a) CH_2F_2 **(b)** $AsCl_5$ **(c)** NH_4^+ **(d)** BF_4^-

18.

Use VSEPR theory to predict the molecular shapes of NO_2^+ and NO_2^-.

19.

Draw Lewis structures for the following molecules and ions, and use VSEPR theory to predict the molecular shape. Indicate the examples in which the central atom has an expanded octet.

(a) XeI_2 **(b)** PF_6^- **(c)** AsF_3 **(d)** AlF_4^-

20.

Given the general formula and the shape of the molecule or ion, suggest possible elements that could be the central atom, X, in each of the following:

(a) XF_3^+ (trigonal pyramidal) **(c)** XF_3 (T-shaped)

(b) XF_4^+ (tetrahedral)