INDEPENDENT LEARNING CENTRE
CENTRE D'ÉTUDES INDÉPENDANTES

Chemistry SCH4U-C
Practice Test

Time: 2 hours
Total Marks: 103 **Final Test Score: ____ ÷ 103 × 100 = ____%**

Instructions

- There is a label attached to this page. Compare the course code on the label with the course code printed on the Final Test to make sure that they are the same. Tell the Final Test supervisor **immediately** if they are not the same.
- The Final Test pages are numbered 1 through 20. Check to see that all 20 pages are there. Tell the Final Test supervisor **immediately** if there are any pages missing.
- You may use a non-marked dictionary during the Final Test if one is available at the test site. You may not use any other books or notes.
- You will need a pen and a scientific calculator.
- There is a periodic table of the elements included at the beginning to assist you with the chemistry questions.
- You must write your answers in the space provided.
- This test has three parts. A breakdown of the marks and the approximate time needed for each part is given below. Look over the entire test before you begin. Manage your time carefully and leave some time at the end to review your work.

Part	Activity	Marks	Time (minutes)
	Preview		5
A	Structure of Matter and Organic Chemistry	25	30
B	Oxidation Reduction Reactions, Electrochemistry, and Energy	39	40
C	Rates of Reaction and Equilibrium Applications	39	40
	Review		5
Total		**103**	**120**

- At the end of the test, return this test paper and all of your work (including drafts) to the Final Test supervisor.
- Please note that, for security reasons, marked tests are not returned to students.
- Mathematical solutions require a proper problem-solving format. Final answers must include the correct units and the correct number of significant figures.

Formula Sheet and Periodic Table of the Elements

Formula Sheet

You will need a scientific calculator that is able to do exponents, logarithms, square roots, and cube roots.

$$n = \frac{m}{M}$$

$$c = \frac{n}{V}$$

$$q = mc\Delta T$$

$$\Delta H = \Sigma\Delta H_{f(products)} - \Sigma\Delta H_{f(reactants)}$$

$$K_{eq} = \frac{[products]}{[reactants]}$$

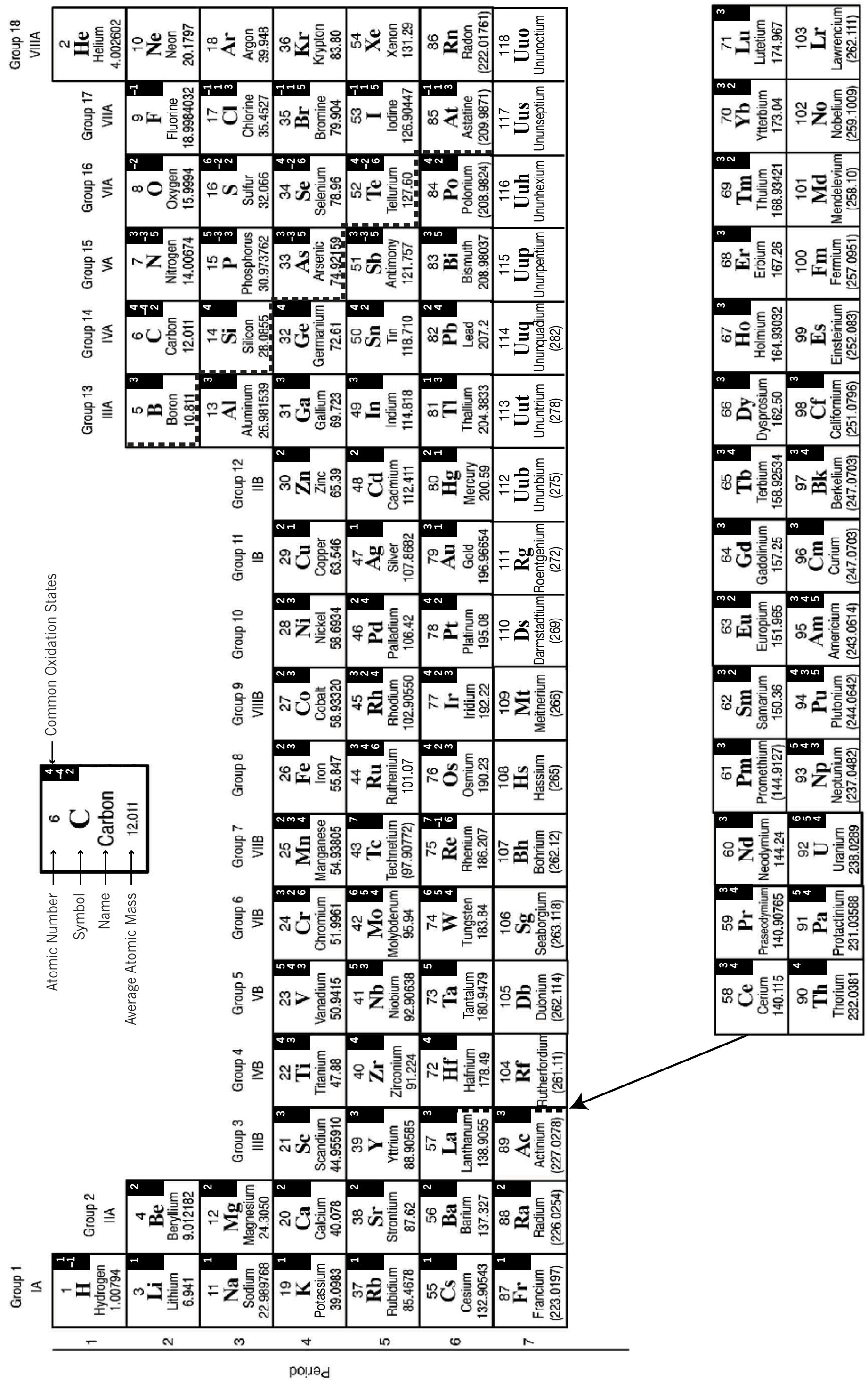
$$(\text{for } \mathrm{A}_x\mathrm{B}_y) \quad K_{sp} = [\mathrm{A}^+]^x[\mathrm{B}^-]^y$$

$$K_a = \frac{[\mathrm{H_3O^+}][\mathrm{A^-}]}{[\mathrm{HA}]}$$

$$K_b = \frac{[\mathrm{HB^+}][\mathrm{OH^-}]}{[\mathrm{B}]}$$

$$\mathrm{pH} = -\log[\mathrm{H_3O^+}]$$

Periodic Table



Part A: Structure of Matter and Organic Chemistry 25 marks *(approximate time: 30 minutes)*

1. Two of the lines in the hydrogen emission spectrum are coloured violet and blue-green. Violet light is known to have more energy than blue-green light. How did Niels Bohr explain the origin of these two lines? **(4 marks)**

Electrons are found in very specific orbits around the atom. **(1 mark)** When external energy is applied to hydrogen's electron, the electrons absorb only the energy corresponding to the difference between orbits or energy levels. **(1 mark)** When an electron absorbs energy, it jumps to a higher orbit. **(1 mark)** When it falls back, it gives off different colours of light. The electron that produced violet light fell farther because violet corresponds to more energy than blue-green. **(1 mark)**

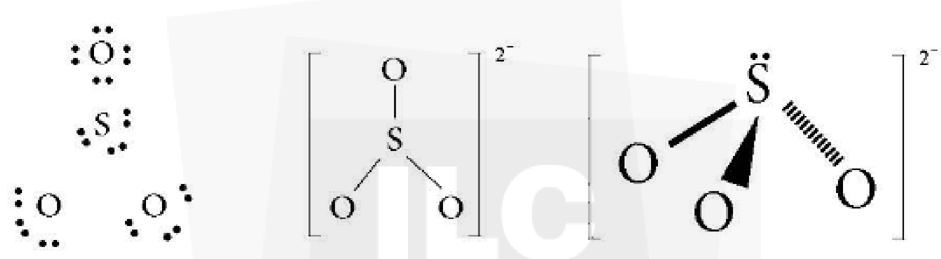
2. Write the electronic configuration for an atom of potassium, which is element 19 on the periodic table. **(2 marks)**

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$ **(2 marks)**

Hint: For a review of how this answer is calculated, see Lesson 1, under the heading "Electronic Configurations" for the steps involved in writing electronic configurations.

3. **a)** Draw the Lewis structure (that is, the electron dot diagram) and structural formula for the sulfite ion, SO_3^{2-}. **(4 marks)**

The correct Lewis structure and structural formula are shown below. **(4 marks: 3 marks for the correct electron arrangement around each atom in the Lewis structure and 1 mark for the correct structural formula)**



b) Predict the shape of the sulfite ion. **(1 mark)**

Since there are three bonds and a lone pair around the central sulfur atom, the shape of this ion is pyramidal. **(1 mark)**

c) Is the sulfite ion polar or non-polar? Justify your prediction. **(3 marks)**

(3 marks total)

The sulfite ion is polar. **(1 mark)** Because the polyatomic ion is pyramidal, the central atom, sulfur, is at the apex of the pyramid. The bond dipoles all face in a general downward direction, away from the lone pair of electrons. Since the dipoles don't cancel each other out, sulfite is polar. **(2 marks)**

4. Methanol and ethane have nearly the same molar mass. Refer to intermolecular forces to explain why there is such a large difference in their boiling points.

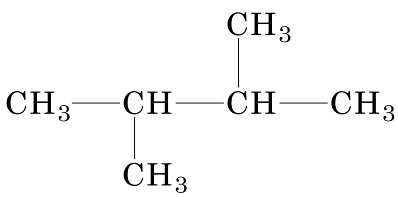
Compound	Formula	Boiling point (°C)
Ethane	C_2H_6	–89
Methanol	CH_3OH	65

(3 marks)

Methanol is a polar molecule, while ethane is non-polar. **(1 mark)** The higher boiling point of methanol is a result of hydrogen bonding between its molecules. **(1 mark)** Ethane has London forces between its molecules, which are considerably weaker than hydrogen bonds. **(1 mark)**

5. Name these compounds. **(4 marks total)**

a)

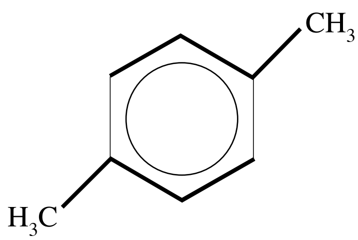


To name this molecule, note that:

- The longest carbon chain is a butane.
- Methyl groups are located at the second and third carbon atoms

Therefore, this molecule is called 2,3-dimethylbutane.

b)

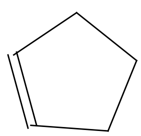


To name this molecule, note that:

- The ring structure is benzene.
- One methyl group is assigned position #1. The other is located on the fourth carbon from carbon #1 in the ring.

Therefore, this molecule is called 1,4-dimethylbenzene.

c)

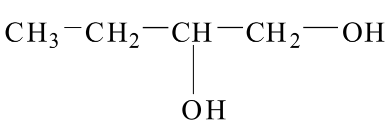


To name this molecule, note that:

- The ring structure is cyclopentene.
- The "-ene" ending indicates that the ring contains a double bond.

Therefore, this molecule is called cyclopentene.

d)



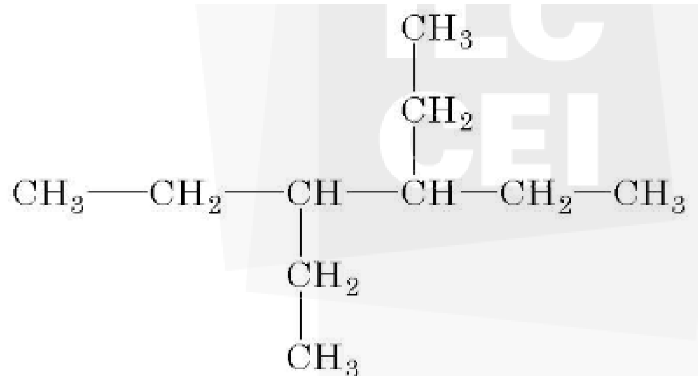
To name this molecule, note that:

- The longest carbon chain is a butane.
- Hydroxyl groups are located at the first and second carbon atoms.
- The presence of hydroxyls makes this an alcohol, hence it has an "-ol" ending.

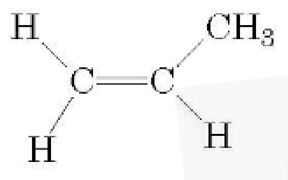
Therefore this molecule is called 1,2-butandiol.

6. Draw the structural formula for these compounds. **(4 marks total)**

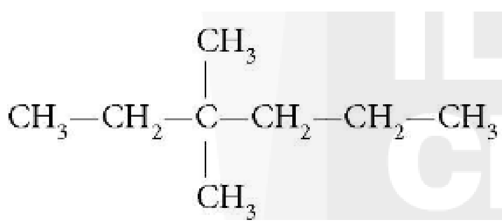
a) 3,4-diethylhexane



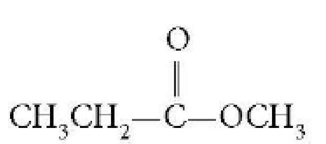
b) propene



c) 3,3-dimethylhexane



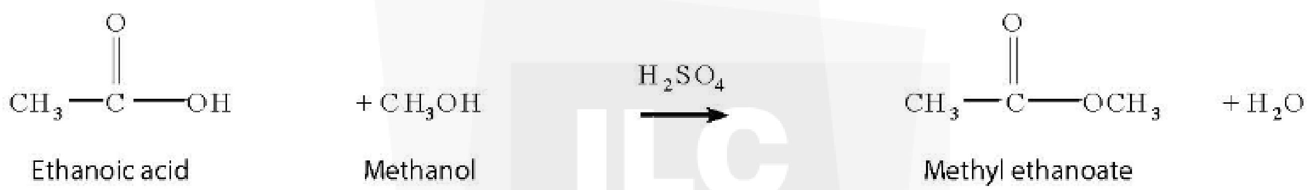
d) methyl propanoate



Part B: Oxidation Reduction Reactions, Electrochemistry, and Energy 39 marks
(approximate time: 40 minutes)

7. Write a balanced equation for each of the following reactions.

a) The reaction of methanol with ethanoic acid in the presence of a sulfuric acid catalyst. Use structural formulas for all organic compounds in your equation. **(3 marks)**



b) The complete combustion of ethanol. Use molecular formulas for each chemical. **(3 marks)**

$C_2H_6O_{(g)} + 3O_{2(g)} \rightarrow 2CO_{2(g)} + 3H_2O_{(g)}$

Note to student: Although it is indicated here, the physical state of each chemical (the "g") is not required to receive full marks for this question.

8. The following table shows the boiling points of alkanes and the corresponding alkyl chlorides.

Alkane	Boiling point (°C)
CH_4	–164
C_2H_6	–89
C_3H_8	–42

Alkyl chloride	Boiling point (°C)
CH_3Cl	–24
C_2H_5Cl	12
C_3H_7Cl	14

a) What effect does the size of the molecule have on the boiling point for both groups of compounds? **(1 mark)**

Boiling point increases as the molecular size increases. **(1 mark)**

b) What effect does the presence of a chlorine atom have on the boiling point of the corresponding alkyl chlorides? **(1 mark)**

The presence of a chlorine atom results in a higher boiling point for the corresponding alkyl chlorides. **(1 mark)**

c) Identify the type of intermolecular forces that occurs in both groups of compounds. Use the difference in forces to explain the difference in boiling points. **(3 marks)**

The intermolecular bonding in alkanes is London forces, while the intermolecular bonding in alkyl chlorides is dipole–dipole attractions. **(1 mark)** This difference occurs because alkanes are non-polar, while the given alkyl chlorides are polar. **(1 mark)**

Since dipole–dipole attractions are stronger than London forces, alkyl halides have higher boiling points. **(1 mark)**

9. a) Chloroethene is the monomer used to manufacture the polymer called polyvinyl chloride. Distinguish between the terms "monomer" and "polymer." **(2 marks)**

A polymer is a long molecule that has a regular repeating unit. **(1 mark)**
A monomer is a small molecule that is used or polymerized to produce the polymer. **(1 mark)**

b) Write a chemical equation showing the synthesis of polyvinyl chloride. **(2 marks)**

c) Why is this reaction an example of addition polymerization? **(1 mark)**

Each chloroethene molecule is added to the forming chain, much like a pearl is added to a necklace. **(1 mark)**

10. Use half-reactions to write a balanced chemical equation for the following net ionic equation. Assume that the reaction is occurring in water. Show each step in the process. **(6 marks)**

$Cr_2O_7^{2-} + NO \rightarrow Cr^{3+} + NO_3^-$

i) The half-reactions:

$Cr_2O_7^{2-} \rightarrow 2Cr^{3+}$

$NO \rightarrow NO_3^-$ **(1 mark)**

ii) Balance oxygen and hydrogen:

$14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$

$2H_2O + NO \rightarrow NO_3^- + 4H^+$ **(1 mark)**

iii) Balance charges by adding electrons:

$14H^+ + Cr_2O_7^{2-} + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

$2H_2O + NO \rightarrow NO_3^- + 4H^+ + 3e^-$ **(1 mark)**

iv) Electrons gained = electrons lost (multiply second equation by 2)

$14H^+ + Cr_2O_7^{2-} + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

$4H_2O + 2NO \rightarrow 2NO_3^- + 8H^+ + 6e^-$ **(1 mark)**

v) Combine and simplify:

$14H^+$	+	$Cr_2O_7^{2-}$	+	$6e^-$	$\rightarrow$	$2Cr^{3+}$	+	$7H_2O$		
		$4H_2O$	+	$2NO$	$\rightarrow$	$2NO_3^-$	+	$8H^+$	+	$6e^-$
$6H^+$	+	$Cr_2O_7^{2-}$	+	$2NO$	$\rightarrow$	$2Cr^{3+}$	+	$2NO_3^-$	+	$3H_2O$ (Final equation)

(2 marks)

11. Consider the following reduction half-cell reactions and potentials:

$Ag^{+}_{(aq)} + e^{-} \rightarrow Ag_{(s)}$ $E° = +0.80\ V$

$Zn^{2+}_{(aq)} + 2e^{-} \rightarrow Zn_{(s)}$ $E° = -0.76\ V$

a) In the galvanic cell that uses these half-cells, predict which reaction will occur at the anode and which will occur at the cathode, and state the type of reaction each will be. **(2 marks)**

$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e$ $E° = 0.76\ V$ Oxidation (occurs at the anode) **(1 mark)**

$Ag^{+}_{(aq)} + 1e \rightarrow Ag_{(s)}$ $E° = 0.80\ V$ Reduction (occurs at the cathode) **(1 mark)**

b) Use this information to determine the cell potential if this redox reaction were occurring in a galvanic cell. **(1 mark)**

Adding the two half-reactions in the redox:

$Zn_{(s)} + Ag^{+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Ag_{(s)}$ $E°_{cell} = 0.76\ V + 0.80\ V = 1.56\ V$ **(1 mark)**

12. These are two half-reactions for a hydrogen fuel cell:

Anode reaction: $2H_2 \rightarrow 4H^{+} + 4e^{-}$

Cathode reaction: $O_2 + 4H^{+} + 4e^{-} \rightarrow 2H_2O$

a) Write the overall cell reaction for the hydrogen fuel cell. **(1 mark)**

Anode reaction:					$2H_2$	$\rightarrow$	$4H^{+}$	+	$4e^{-}$
Cathode reaction:	O_2	+	$4H^{+}$	+	$4e^{-}$	$\rightarrow$	$2H_2O$		
Overall reaction:			$2H_2$	+	O_2	$\rightarrow$	$2H_2O$		

b) Refer to your answer in (a) to explain why running a car on hydrogen fuel cells would be better for the environment than running it on gasoline. **(2 marks)**

The hydrogen fuel cell is better for the environment because its only waste product is water. **(1 mark)** In contrast, the gasoline engine produces a number of products that are harmful to the environment, such as greenhouse gases (carbon dioxide, carbon monoxide) and soot. **(1 mark)**

13. A student heated 100.0 mL of water in an aluminum container using an alcohol burner filled with ethanol (C_2H_5OH). The mass of the aluminum container was 42.5 g. The initial temperature of the water was 17.5°C and the final temperature was 26.3°C. The initial mass of the alcohol burner was 113.45 g and the final mass was 112.97 g. Calculate the molar heat of the combustion of ethanol. The specific heat capacities are c_{Al} = 0.900 J/g°C and c_{H_2O} = 4.18 J/g°C. **(6 marks)**

It is given that:

m_{H_2O} = 100.0 mL = 100.0 g m_{Al} = 42.5 g

c_{H_2O} = 4.18 J/g°C c_{Al} = 0.900 J/g°C

Two substances—the aluminum container and the water—are gaining heat from the alcohol burner. You can easily calculate the temperature change as follows:

ΔT_{H_2O} = 26.3°C – 17.5°C = 8.8°C ΔT_{Al} = 8.8°C **(1 mark)**

Now, you must calculate q, the quantity of heat lost or gained, for the water and the aluminum.

$q = mc\Delta T$

q_{H_2O} = 100.0 g × 4.18 J/g°C × 8.8°C = 3678.4 J = 3.7 kJ **(½ mark)**

q_{Al} = 42.5 g × 0.900 J/g°C × 8.8°C = 336.6 J = 0.34 kJ **(½ mark)**

Add these amounts to get the total quantity of heat lost or gained.

q_{total} = 3.7 + 0.34 = 4.0 kJ **(½ mark)**

Find the amount of ethanol burned, which is the difference between the initial mass and the final mass of the burner.

Mass of ethanol burned = 113.45 g – 112.97 g = 0.48 g (**½ mark**)

The molar mass (M) of ethanol (C_2H_5OH) can be found by adding the atomic masses of all its atoms, as found on the periodic table.

$M = 2(12.01) + 5(1.01) + 16 + 1.01 = 46.08$

Use the mass (m) and the molar mass (M) to calculate the number of moles of ethanol.

$$n = \frac{m}{M} = \frac{0.48\ \text{g}}{46.08\ \text{g/mol}} = 0.0104\ \text{mol} = 0.010\ \text{mol}$$

The moles of ethanol and the total quantity of heat gained can be used in the molar heat of combustion formula, as follows:

$$\Delta H = -\frac{q}{n} = -\frac{4.0\ \text{kJ}}{0.010\ \text{mol}} = -4.0 \times 10^2$$

ΔH would be negative since this was an exothermic reaction. **(2 marks)**

The molar heat of combustion of ethanol is -4.0×10^2 kJ/mol. **(1 mark for final statement, units, and significant figures)**

14. Butane (C_4H_{10}) has a ΔH of –2871 kJ/mol when it undergoes complete combustion into carbon dioxide and liquid water. Given the following ΔH_f values:

Carbon dioxide = –393.5 kJ/mol

Oxygen gas = 0 kJ/mol

Liquid water = –285.8 kJ/mol

a) Write the balanced thermochemical equation for the complete combustion of 1 mol of butane. **(2 marks)**

To write the balanced equation, start by balancing the carbon atoms and the hydrogen atoms. There are four carbon atoms on the reactant side, so the coefficient in front of the carbon dioxide on the products side is 4. There are 10 hydrogen atoms in the reactants, so the coefficient for water becomes 5 (5 × 2 atoms in each water molecule = 10). Finally, balance the oxygen. There are now 13 oxygen atoms on the product side, so the coefficient on the reactants side for oxygen gas becomes $\frac{13}{2}$ or 6.5 (6.5 × 2 atoms in each oxygen molecule = 13).

$$C_4H_{10(l)} + \frac{13}{2}O_{2(g)} \rightarrow 4CO_{2(g)} + 5H_2O_{(l)} + 2871\ \text{kJ}$$

(2 marks total: award ½ mark for each correct coefficient and ½ mark for the correct formulas of reactants and products, keeping in mind that a coefficient of 1 is assumed. Students could write the thermochemical equation either of two ways; with the energy term as a product, as shown, or using the "ΔH notation" to the right of the balanced equation. Award ½ mark for using the energy term correctly.)

b) Calculate ΔH_f of butane in kJ/mol. **(3 marks)**

Substitute the values provided for the ΔH_f for water, carbon dioxide, and oxygen into the equation. You are given the ΔH for butane, so you can substitute this as well. You are calculating the ΔH_f of butane, one of the reactants, so you represent this by using x. Once you have made all of the substitutions, solve for x. **(3 marks total)**

$$\Delta H = \Sigma\Delta H_{f(products)} - \Sigma\Delta H_{f(reactants)}$$

$$-2871\text{ kJ} = [4\text{ mol}(-393.5\text{ kJ/mol}) + 5\text{ mol}(-285.8\text{ kJ/mol})] - [\frac{13}{2}\text{mol}(0\text{ kJ/mol}) + x]$$

$$-2871\text{ kJ} = -3003\text{ kJ} - x$$

$$x = -3003\text{ kJ} + 2871\text{ kJ} = -132\text{ kJ}$$

(2 marks)

The ΔH_f of butane is –132 kJ/mol. **(1 mark for final statement, units, and significant figures)**

Part C: Rates of Reaction and Equilibrium Applications 39 marks *(approximate time: 40 minutes)*

15. The balanced chemical equation for the reaction between iodine vapour and fluorine gas is:

$$3I_{2(g)} + 6F_{2(g)} \rightarrow 2IF_{5(g)} + I_4F_{2(g)}$$

A chemist measures the rate at which iodine vapour is consumed to be 4.5×10^{-4} mol/(L•s). At what rate is $F_{2(g)}$ consumed? **(2 marks)**

(2 marks total; ½ mark for final statement)

The rates of reaction are proportional to the coefficients in the balanced equation. Set up the ratio using the coefficients of the chemicals from the balanced equation, in this case, $I_{2(g)}$ and $F_{2(g)}$. This ratio is equivalent to the ratio of rates, one of which is given, while the other you are required to calculate. The unknown is represented by x.

$$\frac{3\text{ mol } I_2}{6\text{ mol } F_2} = \frac{4.5\times10^{-4}\text{ mol/L•s}}{x\text{ mol/L•s}}$$

$$3x = 6(4.5\times10^{-4})$$

$$x = 9.0\times10^{-4}\text{ mol/L•s}$$

$F_{2(g)}$ is being consumed at the rate of 9.0×10^{-4} mol/L•s.

16. Use the collision theory to explain the following. **(2 marks total)**

a) Chopped vegetables will cook faster than whole vegetables.

The chopped vegetables have a greater surface area that allows more collisions to occur, so the cooking reaction happens faster. **(1 mark)**

b) When a piece of magnesium is held in a flame, it ignites.

Magnesium ignites when held in a flame because heating increases the amount of (and energy of) collisions between magnesium atoms and oxygen molecules. **(1 mark)**

17. Use the concentration-rate data presented in the following table to answer the following questions about the hypothetical reaction 2A + 3B → 2C.

[A] in mol/L	[B] in mol/L	Rate in mol/L•s
0.100	0.100	3.0×10^{-4}
0.300	0.100	9.0×10^{-4}
0.300	0.200	3.6×10^{-3}
0.200	0.300	**?**

a) What is the order of reaction, with regard to each reactant? **(2 marks)**

You need to compare lines 1 and 2: [A] triples, and the rate triples. Because the change in rate is identical to the change in concentration, the order of reaction for reactant A is 1.

You need to compare lines 2 and 3 because [B] changes, but [A] does not: [B] doubles, and the rate quadruples. Because the change in rate is squared, when compared to the change in concentration, the order of reaction for reactant B is 2. **(2 marks)**

b) Write the rate law equation for this reaction. **(1 mark)**

The rate law equation is rate = $k[A][B]^2$ (because the order of reactions become the exponents in the rate law equation). **(1 mark)**

c) Calculate the rate indicated by the "?" **(2 marks)**

You may choose to solve for the value of *k* in the rate law equation and then use the complete rate equation to solve for rate in line 4. The rate in line 4 can also be determined knowing how each reactant affects rate (order of each). When [A] doubles, the rate doubles; when [B] triples, the rate increases by a factor of 9. The overall rate increase is 9 × 2, so the reaction rate increases by a factor of 18.

Therefore, in the final row, rate = $18(3.0 \times 10^{-4}) = 5.4 \times 10^{-3}$ mol/L•s. **(2 marks)**

18. The following equilibrium reaction has a K_{eq} value of 2.5×10^{-6}:

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

Calculate the equilibrium concentrations of all gases if 0.800 mol of $CO_{2(g)}$ is initially placed in a 2.00 L reaction container. **(5 marks)**

$$K_{eq} = \frac{[CO]^2[O_2]}{[CO_2]^2}$$

$$c = \frac{n}{V} = \frac{0.800 \text{ mol}}{2.00 \text{ L}} = 0.400 \text{ mol/L}$$ **(½ mark)**

You need to use the concentration, not the number of moles, in the ICE chart.

ICE chart (1 mark)

ICE	$[CO_2]$ in mol/L	[CO] in mol/L	$[O_2]$ in mol/L
Initial	0.400	0	0
Change	$-2x$	$+2x$	$+x$
Equilibrium	$0.400 - 2x$	$2x$	x
Revised equilibrium	0.400	$2x$	x

Now you apply the 100 rule. If you calculate a value that is greater than 100, then you do not need to use the quadratic equation.

Substitute the revised equilibrium values into the K_{eq} expression and solve for *x*.

$$\frac{[initial]}{K_{eq}} = \frac{0.400 \text{ mol/L}}{2.5 \times 10^{-6}} = 160\,000 > 100$$

$$2.5 \times 10^{-6} = \frac{(2x)^2(x)}{(0.400)^2}$$

$$4x^3 = 2.5 \times 10^{-6}(0.160) = 4.0 \times 10^{-7}$$

(2 marks)

$$x^3 = 1.0 \times 10^{-7}$$

$$x = \sqrt[3]{1.0 \times 10^{-7}} = 4.6 \times 10^{-3}$$

At equilibrium:

$[CO_2]$ = 0.400 mol/L (This value comes directly from the ICE chart.)

$[CO] = 2x = 2(4.6 \times 10^{-3}) = 9.2 \times 10^{-3}$ mol/L **(1 mark)**

$[O_2] = x = 4.6 \times 10^{-3}$ mol/L **(½ mark)**

19. Le Châtelier's principle states that when a system at equilibrium is affected by some kind of stress and loses equilibrium, the system will return to equilibrium by shifting the reaction either to the right or left (forwards or backwards) to minimize the effect of the stress. Use Le Châtelier's principle to describe and explain the shifts caused by the listed stresses placed on the following equilibrium reaction. **(5 marks total)**

$3Cl_{2(g)} + CS_{2(g)} \rightleftharpoons CCl_{4(g)} + S_2Cl_{2(g)}$ $\Delta H = -465$ kJ

a) Adding $CCl_{4(g)}$

The reaction shifts to the left. Shifting to the left reduces the concentration of excess product caused by adding CCl_4. **(1 mark)**

b) Adding $Cl_{2(g)}$

The reaction shifts to the right. Shifting to the right reduces the concentration of excess reactant caused by adding Cl_2. **(1 mark)**

c) Adding a catalyst

There is no shift; the catalyst equally increases the rates of both the forward and reverse reactions. **(1 mark)**

d) Decreasing the volume

The reaction shifts to the right. Shifting to the right decreases volume and increases pressure. This happens because the side with the fewest gas particles has a lower pressure. **(1 mark)**

e) Increasing the temperature

The reaction shifts to the left; heat is a product because the reaction is exothermic. The reaction must shift to the left to reduce excess product. **(1 mark)**

20. The K_{sp} for MgF_2 is 6.4×10^{-9}. What mass of MgF_2 will dissolve in 500 mL of distilled water? **(5 marks)**

(5 marks total)

$MgF_{2(s)} \rightleftharpoons Mg^{2+}{}_{(aq)} + 2F^{-}{}_{(aq)}$

When magnesium fluoride creates a saturated solution, an equilibrium is set up between the undissolved solid and the ions in solution. For every one mole of $Mg^{2+}{}_{(aq)}$, there are two moles of $F^{-}{}_{(aq)}$. That is why you use x and $2x$ in the ICE chart.

$K_{sp} = [Mg^{2+}][F^{-}]^2 = 6.4 \times 10^{-9}$

ICE chart (1 mark)

ICE	Molar solubility in mol/L	$[Mg^{2+}]$ in mol/L	$[F^{-}]$ in mol/L
Initial	x	0	0
Change	________	$+x$	$+2x$
Equilibrium	________	x	$2x$

Substitute equilibrium values from the ICE chart into the K_{sp} expression and solve for x, which is equal to the molar solubility or concentration of the saturated solution.

$$6.4 \times 10^{-9} = x(2x)^2 = 4x^3$$

$$x^3 = \frac{6.4 \times 10^{-9}}{4} = 1.6 \times 10^{-9}$$ **(2 marks)**

$$x = \sqrt[3]{1.6 \times 10^{-9}} = 1.2 \times 10^{-3}$$

The molar solubility (or concentration) of MgF_2 is 1.2×10^{-3} mol/L.

Determine the number of moles of MgF_2 found in 500 mL. The 500 mL has to be converted to L before substitution.

$n = cV = 1.2 \times 10^{-3}$ mol/L $\times$ 0.500 L $= 6.0 \times 10^{-4}$ mol

From the number of moles, determine the mass. The molar mass (M) of MgF_2 can be found by adding the atomic masses of all its atoms, as found on the periodic table.

Molar mass is (1 × 24.3 g/mol) + (2 × 19.0 g/mol) = 62.3 g/mol.

$m = nM = 6.0 \times 10^{-4}$ mol × 62.3 g/mol = 0.037 g **(1½ marks)**

Therefore, 0.037 g of MgF_2 will dissolve in 500.0 mL of distilled water. **(½ mark for the final statement and correct significant figures and units)**

21. **a)** What is the pH of a strong acid with [HA] = 0.125 mol/L? **(1 mark)**

A strong acid is fully ionized. There are no molecules of HA present. Therefore, $[HA] = [H_3O^+] = 0.125$ mol/L.

$pH = -\log[H_3O^+] = -\log(0.125) = 0.903$ **(1 mark)**

b) What is the pH of a weak acid with [HA] = 0.125 mol/L and $K_a = 5.4 \times 10^{-7}$? **(4 marks)**

A weak acid is not fully ionized. $[H_3O^+]$ will have to be calculated using the K_a expression.

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$$

ICE chart (1 mark)

ICE	[HA] in mol/L	$[H_3O^+]$ in mol/L	$[A^-]$ in mol/L
Initial	0.125	0	0
Change	$-x$	$+x$	$+x$
Equilibrium	$0.125 - x$	x	x
Revised equilibrium	0.125	x	x

Now, apply the 100 rule. If you calculate a value that is greater than 100, then you do not need to use the quadratic equation.

Substitute the revised equilibrium values into the K_a expression and solve for x.

$$\frac{[initial]}{K_a} = \frac{0.125 \text{ mol/L}}{5.4 \times 10^{-7}} = 231\,481 > 100$$

$$5.4\times10^{-7}=\frac{(x)(x)}{0.125}=\frac{x^2}{0.125}$$

$x^2=(5.4\times10^{-7})(0.125)=6.8\times10^{-8}$ **(1½ marks)**

$x=\sqrt{6.8\times10^{-8}}=2.6\times10^{-4}$

$[H_3O^+] = x = 2.6 \times 10^{-4}$ mol/L

pH = $-\log(2.6 \times 10^{-4}) = 3.6$ **(1 mark)**

The pH of the 0.125 mol/L weak acid is 3.6. (½ **mark for final statement, and correct significant figures and units)**

22. During a titration, 28.0 mL of 0.150 mol/L $H_2SO_{4(aq)}$ neutralizes 35.0 mL of $NH_4OH_{(aq)}$. The K_a for H_2SO_4 is very large. The K_b for NH_4OH is 1.8×10^{-5}.

a) Write the balanced equation for the neutralization reaction. **(1 mark)**

$H_2SO_{4(aq)} + 2NH_4OH_{(aq)} \rightarrow (NH_4)_2SO_{4(aq)} + 2H_2O_{(l)}$ **(1 mark)**

A neutralization reaction is: acid + base → salt + water. The water is formed from the H^+ of the acid and the OH^- of the base. The positive ion of the base combines with the negative ion of the acid to create the salt.

b) Calculate the concentration of the $NH_4OH_{(aq)}$. **(2 marks)**

First calculate the number of moles of acid.

$n_a = c_aV_a = 0.150 \text{ mol/L} \times 0.028 \text{ L} = 0.0042 \text{ mol}$

Then calculate the number of moles of base, using the mole ratio from the balanced equation.

$$\frac{1 \text{ mol } H_2SO_4}{2 \text{ mol } NH_4OH}=\frac{0.0042 \text{ mol}}{x \text{ mol}}$$

$x = 0.0084$ mol

$$c=\frac{n}{V}=\frac{0.0084 \text{ mol}}{0.035 \text{ L}}=0.24 \text{ mol/L}$$

Finally, calculate the concentration of the base.

The concentration of NH_4OH is 0.24 mol/L. **(2 marks)**

c) From the table below, select an indicator to use in a titration. Explain your selection. **(2 marks)**

Indicator	Range of pH of colour change
Bromocresol purple	5.2 – 6.8
Litmus	6.0 – 8.0
m-cresol purple	7.6 – 9.2

The best indicator to use is bromocresol purple. The strong acid and weak base create an acidic salt. The bromocresol purple is the only indicator in the list that changes colour in the acidic pH range. **(2 marks: 1 mark for the indicator and 1 mark for a well-stated reason)**

23. A buffer is created by adding 0.35 mol of sodium benzoate to 1.00 L of a 0.40 mol/L benzoic acid solution. The K_a of benzoic acid is 6.3×10^{-5}. What is the pH of the buffer? **(5 marks)**

(5 marks total)

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$$

ICE chart (2 marks)

The benzoate ion from the sodium benzoate is the conjugate base of benzoic acid. The buffer already contains both the weak acid and its conjugate base (line 1 of the ICE chart). The weak acid then partially ionizes (line 2 of the ICE chart) and reaches equilibrium (line 3).

ICE	[*HA*] in mol/L	[H_3O^+] in mol/L	[A^-] in mol/L
Initial	0.40	0	0.35
Change	$-x$	$+x$	$+x$
Equilibrium	$0.40-x$	x	$0.35+x$
Revised equilibrium	0.40	x	0.35

Now, apply the 100 rule. If you calculate a value that is greater than 100, then you do not need to use the quadratic equation.

Substitute the revised equilibrium values into the K_a expression and solve for x.

$$\frac{[initial]}{K_a}=\frac{0.40\ \text{mol/L}}{6.3\times10^{-5}}=6349\text{, which is greater than 100}$$

Substitute known values into the K_a expression.

$$6.3\times10^{-5}=\frac{(x)(0.35)}{0.40}$$

(2 marks)

$$x=\frac{(6.3\times10^{-5})(0.40)}{(0.35)}=7.2\times10^{-5}$$