Final Examination Rules

Before you begin your exam:
1. Any student found with an electronic communication device IN THEIR POSSESSION (from the moment they step into the exam room until they pick up their bags after finishing their exam) whether it is used or not WILL BE DISQUALIFIED. (If you do have an electronic communication device, notify an invigilator immediately.)
2. All pencil cases, calculator covers must be below your chair and only I.D. cards, pens, pencils, erasers, calculators and any other allowed course specific materials can be on your desk.
3. Programmable calculators are not permitted. Calculators may not be shared.
4. You may not open the examination booklets, or read examination questions prior to the commencement of the exam. The examination coordinator or his/her representative will announce the beginning and the end of each examination.
5. Write your name and fill out any other required information on the cover page of your exam.

During the exam:
6. You are not allowed under any circumstances to get up during the exam without permission.
Look for your exam number on the SIGN IN sheet that will be brought around by your teacher and sign your name.
7. You are not allowed to leave the gym area during the first hour of the exam.
8. If you need to use the washroom, raise your hand and an invigilator will come see you.
9. You are expected to abide by the rules outlined by the examination proctor or his/her delegate and be aware of College policy regarding cheating and plagiarism.

Once you finished your exam:
10. Raise your hand and stay seated until your teacher will come see you to SIGN OUT next to your exam number.
11. During the last 15 minutes of the exam, no students are allowed to leave.
12. Once the examination coordinator announces the end of the exam, everyone must stay seated, quietly, until all exams are collected.

Final Examination Instructions

1. This exam set consists of 19 questions. Please ensure that you have a complete set.
2. **Answer all questions in the space provided.** No books or extra paper are permitted.
3. In order to obtain full credit, you must show the method used to solve all problems involving calculations and express your answers to the correct number of significant figures.
4. Your attention is drawn to the College policy on cheating.

5. **A Periodic Table and useful data is provided on the last page. You may detach the last sheet.**

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<table>
<thead>
<tr>
<th>Question</th>
<th>Marks</th>
</tr>
</thead>
</table>

Sig. Figs./Units __________ /2

TOTAL________ /100
Q1. a) Name the following compounds.  

i. Mg(ClO$_4$)$_2$ ........................................... magnesium perchlorate  

ii. N$_3$O$_2$ ............................................... trinitrogen dioxide  

iii. SnI$_2$ ...................................................... tin(II) iodide  

iv. NaH ......................................................... sodium hydride  

v. NiHSO$_4$ .................................................. nickel (I) hydrogen sulfate or bisulfate  

b) Write the formula for the following compounds.  

i. aluminum fluoride .............................................................. AlF$_3$  

ii. hydrosulfuric acid ............................................................... H$_2$S$_{(aq)}$  

iii. copper(II) nitride ................................................................. Cu$_2$N$_2$  

iv. nitrous acid .................................................................... HNO$_2$$_{(aq)}$  

v. dichlorine heptoxide ........................................................... Cl$_2$O$_7$  

Q2. Roundup™, a herbicide manufactured by Monsanto, has the formula C$_3$H$_8$NO$_5$P.  
How many molecules are there in a 405.3 g sample of Roundup?  

\[
\text{Molar mass of roundup} = 169.09 \text{ g/mole}  
\]

\[
405.3 \text{ g} / 169.09 \text{g/mole} = 2.397 \text{ moles}  
\]

\[
2.397 \text{ moles} \times 6.022 \times 10^{23} = 1.443 \times 10^{24} \text{ molecules}  
\]
Q3. Classify each of the reactions given below as: precipitation and/or acid-base and/or oxidation-reduction. (2 marks)

<table>
<thead>
<tr>
<th>Chemical Reactions</th>
<th>Classification</th>
</tr>
</thead>
<tbody>
<tr>
<td>i. $2 \text{Na}_3\text{PO}_4 (aq) + 3 \text{Pb(NO}_3)_2 (aq) \rightarrow \text{Pb}_3(\text{PO}_4)_2 (s) + 6 \text{NaNO}_3 (aq)$</td>
<td>Precipitation</td>
</tr>
<tr>
<td>ii. $2 \text{PbS} (s) + 3\text{O}_2 (g) \rightarrow 2 \text{PbO} (s) + 2\text{SO}_2 (g)$</td>
<td>Redox</td>
</tr>
<tr>
<td>iii. $\text{H}_2\text{SO}_4 (aq) + 2 \text{NaOH} (aq) \rightarrow \text{Na}_2\text{SO}_4 (aq) + 2 \text{H}_2\text{O} (l)$</td>
<td>Acid-base</td>
</tr>
<tr>
<td>iv. $2 \text{Cu(SO}_4)_2 (aq) + 4 \text{KI} (aq) \rightarrow 2 \text{CuI} (s) + \text{I}_2 (aq) + 2 \text{K}_2\text{SO}_4 (aq)$</td>
<td>Precipitation and Redox</td>
</tr>
</tbody>
</table>

Q4. a) Assign an oxidation number to each atom of IO$_3^-$.

\[
\text{MnO}_4^- (aq) + \text{IO}_3^- (aq) \rightarrow \text{MnO}_2 (s) + \text{IO}_4^- (aq)
\]

I: +5  O: -2

b) What is the oxidising agent in the reaction above? Answer: \text{MnO}_4^-(1 mark)

c) Balance the above reaction in basic media using the half reaction method. (4 marks)

Identify the oxidation and reduction half equations.

\[
\text{Red. } \frac{1}{2} \text{ eqn. } [4\text{H}^+ (aq) + \text{MnO}_4^- (aq) + 3 e^- \rightarrow \text{MnO}_2 (s) + 2 \text{H}_2\text{O} (l)] \times 2
\]

\[
\text{Ox. } \frac{1}{2} \text{ eqn. } [\text{H}_2\text{O} (l) + \text{IO}_3^- (aq) \rightarrow \text{IO}_4^- (aq) + 2 e^- + 2\text{H}^+ (aq)] \times 3
\]

\[
8 \text{H}^+ (aq) + 2 \text{MnO}_4^- (aq) + 3 \text{IO}_3^- (aq) \rightarrow 2\text{MnO}_2 (s) + 3\text{IO}_4^- (aq) + \text{H}_2\text{O} (l) + 6\text{H}^+ (aq)
\]

- cancel $6\text{H}^+ (aq)$ from both sides of reaction equation, then add $2\text{OH}^- (aq)$ to both sides of the reaction equation to give

\[
2 \text{H}^+ (aq) + 2\text{OH}^- (aq) + 2\text{MnO}_4^- (aq) + 3 \text{IO}_3^- (aq) \rightarrow 2\text{MnO}_2 (s) + 3\text{IO}_4^- (aq) + \text{H}_2\text{O} (l) + 2\text{OH}^- (aq)
\]

\[
\text{H}_2\text{O} (l) + 2 \text{MnO}_4^- (aq) + 3 \text{IO}_3^- (aq) \rightarrow 2\text{MnO}_2 (s) + 3\text{IO}_4^- (aq) + 2\text{OH}^- (aq)
\]
Q5. Qualitative tests on a sample of waste water resulted in a positive test for the presence of Fe\(^{3+}\) ion. A 20.00 ml sample of the waste water was diluted to 100.00 mL with distilled water. Titration of the *diluted* solution required 31.76 mL of a 3.664 \(\times 10^{-3}\) M KOH solution to fully precipitate the Fe\(^{3+}\) as a reddish-brown precipitate.

a) Write the net ionic equation (include states of matter) for the reaction of KOH solution and Fe\(^{3+}\) ion in the waste water. 

\[
\text{Fe}^{3+} (aq) + 3\text{OH}^- (aq) \rightarrow \text{Fe(OH)}_3 (s)
\]

b) Determine the molar concentration of Fe\(^{3+}\) ion in the original 20.00 mL sample of waste water. 

\[
\left[\text{Fe}^{3+}\right] = \frac{0.03176 \text{ L} \times 0.003354 \text{ mol OH}^- \times \frac{1 \text{ mol Fe}^{3+}}{3 \text{ mol OH}^-} \times \frac{1000 \text{ mL}}{20 \text{ mL}}}{0.03176 \text{ L} \times 0.003354 \text{ mol OH}^- \times \frac{1 \text{ mol Fe}^{3+}}{3 \text{ mol OH}^-} \times \frac{1000 \text{ mL}}{20 \text{ mL}}}
\]

\[
\left[\text{Fe}^{3+}\right]_{\text{waste water}} = 0.001939 \text{ M}
\]
Q6. a) Write the complete molecular, complete ionic, and net ionic equations for the reaction that takes place when the following aqueous solutions are mixed. Identify the spectator ion(s). Include all charges and the states of matter for each species.

\[
\text{mercury (II) nitrate and potassium phosphate } \rightarrow ?
\]

C.M.E.: \[3 \text{Hg(NO}_3\text{)}_2 (aq) + 2 \text{K}_3\text{PO}_4 (aq) \rightarrow \text{Hg}_3(\text{PO}_4)_2 (s) + 6 \text{KNO}_3 (aq)\]

C.I.E. \[3\text{Hg}^{2+} (aq) + 6\text{NO}_3^- (aq) + 6\text{K}^+ (aq) + 2\text{PO}_4^{3-} (aq) \rightarrow \text{Hg}_3(\text{PO}_4)_2 (s) + 6\text{K}^+ (aq) + 6\text{NO}_3^- (aq)\]

Spectator Ion(s): \[6\text{K}^+ (aq), 6\text{NO}_3^- (aq)\]

N.I.E.: \[3\text{Hg}^{2+} (aq) + 2\text{PO}_4^{3-} (aq) \rightarrow \text{Hg}_3(\text{PO}_4)_2 (s)\]

b) Write the net ionic equation for the reaction between sodium carbonate and hydrochloric acid.

\[
\text{CO}_3^{2-} (aq) + 2\text{H}^+ (aq) \rightarrow \text{CO}_2 (g) + \text{H}_2\text{O} (l)
\]

Q7. Sublimation of dry ice (frozen CO\textsubscript{2}) is used to create special effects in many concerts. How many litres of CO\textsubscript{2} gas will be formed at a pressure of 775 mmHg and a temperature of 27.0°C if a 2.20 g sample of dry ice is used? Assume that the gas is ideal.

**Solution:**

*Number of moles of CO\textsubscript{2}:*

\[
2.20 \text{ g} / 44.0 \text{ g·mol}^{-1} = 0.0500 \text{ moles CO}_2
\]

*Since ideal gas: PV = nRT → V = nRT/P*

\[
n = 0.0500 \text{ mol} \\
T = 273.15 + 27.0 = 300.15 \text{ K} \\
P = 775 \text{ mmHg} \times 1.00 \text{ atm} / 760 \text{ mmHg} = 1.02 \text{ atm}
\]

\[
V = (0.050 \text{ mol})(0.08206 \text{ L·atm·mol}^{-1}·\text{K}^{-1})(300.15 \text{ K}) / (1.02 \text{ atm})
\]

\[
= 1.208 \text{ L} \text{ of CO}_2 \text{ gas will be formed.}
\]

**Answer:** 1.20(8) L of CO\textsubscript{2}
Q8. a) Determine the empirical formula of the compound that has the following mass percentage:  C = 40.0%,  H = 7.00%,  O = 53.0%.  

In 100 g, we have 40.0 g of C, 7.00 g of H and 53.0 g of O.

The molar mass of C is 12.01 g/mol, the molar mass of H is 1.008 g/mol and the molar mass of O is 15.99 g/mol.

Thus, the number of moles of C is 3.33 mol, the number of moles of H is 6.94 mol and the number of moles of O is 3.31 mol.

The ratio of each element is:

C: \( \frac{3.33}{3.31} = 1 \)
H: \( \frac{6.94}{3.31} = 2.09 = 2 \)
O: \( \frac{3.33}{3.31} = 1 \)

Empirical formula = CH₂O

b) Determine the molecular formula of the acid produced if the molar mass of the compound is between 58 – 66 g/mol.

Molar mass of CH₂O = 12.0 + 2(1.0) + 16.0 = 30.0 g/mol
Molecular formula = C₂H₄O₂ = 60.0 g/mol

(3 marks)

(1 mark)
Q9. a) The total energy of a nitrogen gas laser pulse with a wavelength of 337 nm is 3.87 millijoules.

i. How many photons does the laser pulse contain?

\[ E_{\text{photon}} = \frac{hc}{\lambda} = \frac{6.626 \times 10^{-34} \text{ J.s} \times 2.998 \times 10^8 \text{ m/s}}{337 \times 10^{-9} \text{ m}} = 5.89(5) \times 10^{-19} \text{ J} \]

Number of photons = \( \frac{E_{\text{pulse}}}{E_{\text{photon}}} = \frac{3.87 \times 10^{-3} \text{ J}}{5.89(5) \times 10^{-19} \text{ J}} = 6.56 \times 10^{15} \text{ photons} \)

ii. In which region of the electromagnetic spectrum is the radiation found?

Answer: The laser pulse is in the **ultraviolet** part of the spectrum.

b) An electron in the n = 6 level of a hydrogen atom falls to a lower energy level emitting light of wavelength 93.8 nm. Find the principal energy level to which the electron falls.

\[ \frac{1}{\lambda} = R_H \left( \frac{1}{n_{\text{in}}^2} - \frac{1}{n_{\text{out}}^2} \right) \]

\[ \frac{1}{n_{\text{in}}^2} = \frac{1}{(\lambda R_H)} + \frac{1}{n_{\text{out}}^2} = \left( \frac{1}{93.8 \times 10^{-9} \text{ m} \times 1.0974 \times 10^7 \text{ m}^{-1}} \right) + (1/6^2) \]

\[ = 0.971(5) + 0.2778 \]

\[ = 0.999(3) \]

\[ n_{\text{in}} = \sqrt{1/0.999} \approx 1 \]

The electron falls to the n=1 energy level

c) What is the de Broglie wavelength (in nm) associated with a 2.5 g Ping-pong ball travelling at a speed of 20. m/s?

\[ \lambda = \frac{h}{mv} \quad \text{where} \quad m = 2.5 \text{ g} = 2.5 \times 10^{-3} \text{ kg} \]

\[ = 6.626 \times 10^{-34} \text{ J.s} / (2.5 \times 10^{-3} \text{ kg} \times 20. \text{ m/s}) \]

\[ = 1.3(2) \times 10^{-32} \text{ m} \]

\[ = 1.3 \times 10^{-23} \text{ nm} \]
Q10. a) Which of the four quantum numbers \((n, l, m_l, m_s)\) determine . . . 

i. . . . the energy level of an orbital in a hydrogen atom. Answer: \(n\)
ii. . . . the shape of an orbital. Answer: \(l\)
iii. . . . the size of an orbital Answer: \(n\)
iv. . . . the spatial orientation of an orbital. Answer: \(m_l\)

b) Draw the orbital (box) diagram for the atom with the following electron configuration. Name the element represented by the electron configuration.

\(1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5\)  Element: Cr

\[\begin{array}{ccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccccc

(c) Which of the following sets of quantum numbers are allowed and which are not allowed to specify an electron? For the set of quantum numbers that are incorrect, state what is wrong in each set.

i. \(n = 1, l = 1, m_s = -\frac{1}{2}\) Not allowed, max for \(l\) is \(n-1\)
ii. \(n = 4, l = 3, m_l = -2, m_s = + \frac{1}{2}\) Allowed

(d) Answer the following questions (i. to iii.) for an atom of gallium, Ga. 

i. Write the complete ground state electron configuration.
Answer: \(31\text{Ga}: 1s^22s^22p^63s^23p^64s^23d^{10}4p^1\)

ii. How many valence electrons does Ga have? Answer: 3 valence electrons

iii. How many \(s\) electrons does Ga have with a spin of \(+\frac{1}{2}\)? Answer: 4, \(s\) electrons
Q11. a) Explain why Se\(^{2-}\) has a larger atomic radius than Sr\(^{2+}\).  

The electrons of Sr\(^{2+}\), which are shielded from its nucleus to the same extent as Se\(^{2-}\), are held closer by greater attraction to the extra protons in its nucleus.

b) Answer the following questions (i. – iii.) for the set of atoms B, Al and O:  

i. Predict the order of increasing atomic radius.

<table>
<thead>
<tr>
<th>Smallest Atomic Radius</th>
<th>O</th>
<th>B</th>
<th>Al</th>
</tr>
</thead>
<tbody>
<tr>
<td>Largest Atomic Radius</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

ii. Predict the order of increasing first ionization energy (I.E.\(_1\)).

<table>
<thead>
<tr>
<th>Smallest I.E.(_1)</th>
<th>Al</th>
<th>B</th>
<th>O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Largest I.E.(_1)</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

iii. Order the atoms from the least exothermic electron affinity (EA) to the most exothermic.

<table>
<thead>
<tr>
<th>Least Exothermic EA</th>
<th>Al</th>
<th>B</th>
<th>O</th>
</tr>
</thead>
<tbody>
<tr>
<td>Most Exothermic EA</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

c) Circle the bond in each of the following groups that would be the least polar.

i. Mg-O Be-O Ba-O

ii. Si-F Si-S Si-Cl

d) Predict whether a bond between each of the pairs of atoms below would be nonpolar covalent, polar covalent, or ionic.

i. Cl–C Answer: polar covalent

ii. Br–Br Answer: nonpolar covalent

iii. Mg–N Answer: Ionic
Q12. Use the molecular orbital diagram (right) to answer the following questions.

a) According to MO theory, which of the following molecules would you predict to be paramagnetic? (1 mark)

i. $N_2$
ii. $C_2$
iii. $B_2$
iv. Both $B_2$ and $C_2$
v. Both $B_2$ and $N_2$  Answer: ___ iii. $B_2$ ___

b) Which of the following has the set of molecules correctly arranged in order of increasing bond length? (1 mark)

i. $C_2^+ < C_2 < C_2^-$
ii. $C_2^- < C_2 < C_2^+$
iii. $C_2^- < C_2^+ < C_2$
iv. $C_2 < C_2^+ < C_2^-$
v. $C_2 < C_2^- < C_2^+$  Answer: ___ ii. ___

c) What is the bond order for the molecular anion $C_2^{2-}$? (1 mark)

i. 3
ii. 2.5
iii. 2
iv. 1.5
v. 1  Answer: ___ i. 3 ___
Q13. Piperine, below, is the compound responsible for the spiciness of black pepper.

a) Complete the Lewis structure of piperine below by adding lone-pairs wherever required. (1 mark)

![Lewis structure of piperine](image)

b) Provide the following information for the piperine molecule: (2 marks)

i. Number of carbon atoms that are sp² hybridized: 11

ii. Hybridisation of the nitrogen atom: sp³

iii. Smallest of the labeled bond angles (x, y or z): x (104.5°)

iv. Number of π bonds: 6

Q14. a) Both the coffee-cup calorimeter and the bomb calorimeter can be used to measure the heat involved in a reaction. Which calorimeter measures ΔH and which calorimeter measures ΔE? (1 mark)

i. ΔH is measured by a coffee-cup calorimeter.

ii. ΔE is measured by a bomb calorimeter.

b) In a coffee-cup calorimeter, 1.10 g of NH₄NO₃ (molar mass 80.052 g/mol) is mixed with 76.13 g of water at an initial temperature of 25.00°C. After dissolution of the salt, the final temperature of the calorimeter contents is 23.93°C. Assuming the solution has a heat capacity of 4.18 J/°C·g and assuming no heat is lost to the calorimeter, calculate the enthalpy change for the dissolution of NH₄NO₃ in units of kJ/mol. (3 marks)

\[
(1.10+76.13)\text{g} \times -1.07°C \times 4.18 \text{J/°C} \cdot \text{g} = -345.4(4) \text{J (energy absorbed)}
\]

moles of NH₄NO₃ = 1.10 g / 80.052 g/mol = 0.0137(4) moles

+0.345(4) kJ / 0.0137(4) moles = +25.1 kJ/mol (with correct S.F.)
Q15. a) Iron(III) oxide is reduced by carbon monoxide in a blast furnace through a series of reactions to produce two final products - iron metal (Fe(s)) and one other product. 

Use the equations below to determine the overall balanced equation for the reaction of iron(III) oxide and carbon monoxide to give Fe(s) and the unnamed product and calculate the $\Delta H^\circ_{\text{rxn}}$ for this process.

Reaction 1: $3\text{Fe}_2\text{O}_3(s) + \text{CO}(g) \rightarrow 2\text{Fe}_3\text{O}_4(s) + \text{CO}_2(g)$ $\Delta H^\circ = -48.5 \text{ kJ}$

Reaction 2: $\text{Fe}(s) + \text{CO}_2(g) \rightarrow \text{FeO}(s) + \text{CO}(g)$ $\Delta H^\circ = -11.0 \text{ kJ}$

Reaction 3: $\text{Fe}_3\text{O}_4(s) + \text{CO}(g) \rightarrow 3\text{FeO}(s) + \text{CO}_2(g)$ $\Delta H^\circ = +22.0 \text{ kJ}$

**Answer:**

\[
\begin{align*}
3\text{Fe}_2\text{O}_3(s) + \text{CO}(g) & \rightarrow 2\text{Fe}_3\text{O}_4(s) + \text{CO}_2(g) & \Delta H^\circ = -48.5 \text{ kJ} \\
2\text{Fe}_3\text{O}_4(s) + 2\text{CO}(g) & \rightarrow 6\text{FeO}(s) + 2\text{CO}_2(g) & \Delta H^\circ = +44.0 \text{ kJ} \\
6\text{FeO}(s) + 6\text{CO}(g) & \rightarrow 6\text{Fe}(s) + 6\text{CO}_2(g) & \Delta H^\circ = +66.0 \text{ kJ} \\
\hline
3\text{Fe}_2\text{O}_3(s) + 9\text{CO}(g) & \rightarrow 6\text{Fe}(s) + 9\text{CO}_2(g) & \Delta H^\circ_{\text{rxn}} = 61.5 \text{ kJ}
\end{align*}
\]

**Fe$_2$O$_3$(s) + 3CO(g) → 2Fe(s) + 3CO$_2$(g) $\Delta H_{\text{rxn}}^\circ = 20.5 \text{ kJ}$**

b) i. Write the balanced equation (including states of matter) that represents the standard heat of formation, $\Delta H^\circ_f$, of ammonia gas (NH$_3$). 

\[
\frac{3}{2} \text{H}_2(g) + \frac{1}{2} \text{N}_2(g) \rightarrow \text{NH}_3(g)
\]

ii. Calculate the $\Delta H^\circ_f$ for NH$_3$ (in kJ/mol) using the table of bond energies, right.

\[
\begin{align*}
\Sigma n(\text{bonds broken}) - \Sigma n(\text{bonds formed}) &= 1.5 \times 432 + 0.5 \times 941 - 3 \times 391 \\
&= -54(5) \text{ kJ/mol (0 dp)}
\end{align*}
\]
Q16. Use the following list of ions and molecules to answer the questions below.

\[ \text{CO}_2, \text{ClF}_5, \text{PO}_3^{3-}, \text{PF}_5 \]

a) Which of the molecules or ions, according to the VSEPR model, has 4 electron-pairs around the central atom?

Answer: __PO_3^{3-}____  

(1 mark)

b) Draw the 3D-molecular shape of your answer from a), clearly including lone pairs and bond angles. Name its molecular shape.

![Trigonal pyramid](image)

(2 marks)

c) Which of the molecules or ions is linear?

Answer: __CO_2____  

(1 mark)

d) Draw all of the non-equivalent resonance structures of CO\(_2\). For each resonance structure, assign formal charges to the atoms and circle the structure that contributes most to the resonance hybrid.

![Resonance structures](image)

FC: 0 0 0 -1 0 +1 +1 0 -1

(3 marks)
e) Is the molecule ClF$_5$ polar or nonpolar? Explain your answer by drawing and naming its 3D-molecular structure (including all lone-pair electrons). If the molecule is polar draw the molecule’s dipole moment.

$\text{ClF}_5$ is polar because F is more electronegative than chlorine and since a square pyramidal structure has a lone pair instead of a bond, there is a dipole moment.
Q17. a.) Explain in terms of intermolecular forces why . . .  
   i. NH₃ has a higher boiling point than CH₄.

   Ammonia has H-bonding and methane only vdW forces.

   ii. KCl has a higher melting point than I₂.

   KCl has ionic bond and iodine is covalent and non-polar.

b) The binary hydrogen compounds of Group 4A elements have the following boiling points:

<table>
<thead>
<tr>
<th>Compound:</th>
<th>Boiling point:(°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>CH₄</td>
<td>– 162</td>
</tr>
<tr>
<td>SiH₄</td>
<td>– 112</td>
</tr>
<tr>
<td>GeH₄</td>
<td>– 88</td>
</tr>
<tr>
<td>SnH₄</td>
<td>– 52</td>
</tr>
</tbody>
</table>

Fully explain the increase in boiling points from CH₄ to SnH₄.  

The higher the atomic mass of the central atom the greater its polarisability leading to greater vdW forces between the molecules. The greater the forces between molecules the greater energy/temperature required to break them apart to make a gas.

(2 marks)

(2 marks)

(2 marks)

c) List the most significant intermolecular force that exists between molecules in each of the following species:

i. HF……………………………………………………………………………………………………………………………... H-bonding

ii. CH₃Cl………………………………………………………………………………………………………………………….. dipole-dipole

iii. benzene, C₆H₆……………………………………………………………………………………………………………… van der Waals

iv. CS₂……………………………………………………………………………………………………………………………… van der Waals

(2 marks)
Q18. a) Complete the phase diagram of H₂O below.  

i. Label the axes.

ii. Label the physical state that corresponds to each region.

iii. Circle the triple point.

b) Name all of the physical changes (eg “sublimation”) that would occur to a sample of H₂O if the following procedures were carried out?  

i. Starting at point A, the temperature is raised at constant pressure. 

   Melting then boiling

ii. Starting at point B, the pressure is lowered at constant temperature. 

   Boiling

c) What is the physical significance of the critical point of water?  

At temperatures and pressures beyond the critical point there is no distinct phase change between gas and liquid. For example, at temperatures higher than the critical point temperature, it is impossible to condense a gas into a liquid.
Q19. The following information describes the procedure and results for the reaction of calcium chloride with sodium carbonate to produce calcium carbonate. Answer the following questions based on the data supplied.

**PROCEDURE**

i. Weigh the CaCl$_2$·2H$_2$O and record the mass on the data sheet
ii. Transfer the solid CaCl$_2$·2H$_2$O into a volumetric flask, dissolve the solid with distilled water and fill the flask to the 100.0 mL mark.
iii. Pipette 10.00 mL of this solution into a 150 mL beaker
iv. Record the concentration of the Na$_2$CO$_3$ solution and add, to the same 150 mL beaker a precise volume of Na$_2$CO$_3$ solution. Record this volume on your datasheet.
v. On a clean and dry watch glass, put a filter paper and weigh both of them together.
vi. Use this filter paper to filter the solid CaCO$_3$ obtained, dry your filter paper containing the final product in an oven at 110°C for 20 min and weigh on the watch glass.

**DATA SHEET**

1. Mass of CaCl$_2$·2H$_2$O  
   4.0122 g
2. [Na$_2$CO$_3$], mol·L$^{-1}$  
   0.3330 M
3. Volume of the Na$_2$CO$_3$ solution added  
   11.03 mL
4. Mass of filter paper + watch glass  
   46.8778 g
5. Mass of filter paper + watch glass + dry product (final)  
   47.1001 g

**Molar masses:** Na$_2$CO$_3$ : 105.99 g/mol , CaCO$_3$ : 100.09 g/mol , CaCl$_2$ : 110.98 g/mol

a) Write the balanced molecular equation of this reaction.  

- CaCl$_2$(aq) + Na$_2$CO$_3$(aq) → 2 NaCl(aq) + CaCO$_3$(s)  

b) Calculate the concentration of the initial CaCl$_2$ solution (100.0 mL volumetric flask).  

- 4.0122 g CaCl$_2$·2H$_2$O x (1 mol / 147.01 g) = 2.7292x10$^{-2}$ mole
- 2.7292x10$^{-2}$ mole CaCl$_2$·2H$_2$O / 0.1000 mL = 0.2729 M

The initial concentration of CaCl$_2$ = **0.2729 M**
Q19. continued.

c) Identify the limiting reactant. (2 marks)

Mole of CaCl\textsubscript{2} in the beaker:
0.2729 M x 0.01000 L = 2.729x10\textsuperscript{-3} mole

Mole of Na\textsubscript{2}CO\textsubscript{3} in the beaker:
0.3330 M x 0.01103 L = 3.673x10\textsuperscript{-3} mole

Since the stoichiometric ratio is 1 CaCl\textsubscript{2} for 1 Na\textsubscript{2}CO\textsubscript{3} then **CaCl\textsubscript{2} is the limiting reactant.**

d) Calculate the theoretical yield (in grams). (2 marks)

From the previous calculation:
CaCl\textsubscript{2} = 2.729x10\textsuperscript{-3} mole (limiting reactant)

Since the stoichiometric ratio is 1 CaCl\textsubscript{2} for 1 CaCO\textsubscript{3} then the number of moles of CaCO\textsubscript{3} obtained will be the same as the one of the limiting reactant (assuming 100% yield).

**Theoretical yield:**

\[2.729 \times 10^{-3} \text{ mole CaCl}_2 \times \left(1 \frac{\text{CaCO}_3}{1 \text{ CaCl}_2}\right) \times (100.09 \text{ g/mol}) = 0.2731 \text{ g CaCO}_3\]

e) Calculate the actual yield and the % yield of the reaction. (2 marks)

**Actual yield:**

47.1001 g – 46.8778 g = **0.2223 g CaCO\textsubscript{3}**

**% Yield:**

\[(0.2223 \text{ g} / 0.2731 \text{ g}) \times 100\% = 81.40\%\]
### Periodic Table of the Elements

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<thead>
<tr>
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<th>2A</th>
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</thead>
<tbody>
<tr>
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<td>6</td>
<td>Cs</td>
</tr>
<tr>
<td>7</td>
<td>Fr</td>
</tr>
</tbody>
</table>

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**USEFUL DATA:**

Bohr constant $B = 2.178 \times 10^{-18}$ J

Rydberg constant $R_H = 1.0974 \times 10^7$ m$^{-1}$

Gas constant $R = 0.08206$ L·atm·mol$^{-1}$·K$^{-1} = 8.314$ L·kPa·mol$^{-1}$·K$^{-1}$

Avogadro’s number $N_A = 6.0221 \times 10^{23}$ mol$^{-1}$

Planck’s constant $h = 6.626 \times 10^{-34}$ J·s

Speed of light $c = 2.998 \times 10^8$ m·s$^{-1}$

1 atm = 760 mm Hg = 101.3 kPa = 760 torr

1 J = 1 kg·m$^2$·s$^{-2}$

101.3 J = 1 L·atm

0.0°C = 273.2 K