Print your Name: __________________________________________

Student Number: ______________________________

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INSTRUCTIONS:

This exam set consists of 16 questions. Please ensure that your copy of this examination is complete.

Answer all questions in the space provided.

1. Calculators may not be shared. Programmable calculators are not permitted.
2. 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. This policy will be enforced.
5. A Periodic Table is provided. You may detach the Periodic Table.

USEFUL DATA:

Avogadro’s Number  \( N_A = 6.022 \times 10^{23} \text{ mol}^{-1} \)
Gas Constant \( R = 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1} = 8.314 \text{ L kPa K}^{-1} \text{ mol}^{-1} \)
\( = 8.314 \text{ J K}^{-1} \text{ mol}^{-1} \)
Bohr Orbit Constant \( B = 2.178 \times 10^{-18} \text{ J} \)
Rydberg Constant \( R_H = 1.0974 \times 10^7 \text{ m}^{-1} \)
Planck’s Constant \( h = 6.626 \times 10^{-34} \text{ J s} \)
Speed of light \( c = 2.998 \times 10^8 \text{ m s}^{-1} \)
Mass of an electron \( m_e = 9.11 \times 10^{-31} \text{ kg} \)
Mass of a proton \( m_p = 1.67 \times 10^{-27} \text{ kg} \)
1 atm = 101.3 kPa = 760 mmHg = 760 torr
1 J = 1 kg m² s⁻²

MARK DISTRIBUTION

1. /5
2. /6
3. /5
4. /7
5. /5
6. /6
7. /8
8. /6
9. /5
10. /6
11. /9
12. /7
13. /4
14. /6
15. /8
16. /5

Significant Figures /1
Units /1

TOTAL /100
Question 1

a. Write the formula of each of the following compounds.  

(2.5 marks)

i. Dinitrogen monoxide  \[ \text{N}_2\text{O} \]

ii. Copper (II) sulfate pentahydrate  \[ \text{CuSO}_4\cdot5\text{H}_2\text{O} \]

iii. Tin (IV) oxide  \[ \text{SnO}_2 \]

iv. Phosphorous acid  \[ \text{H}_3\text{PO}_3(\text{aq}) \]

v. Ammonium hydrogen sulfate  \[ \text{NH}_4\text{HSO}_4 \]

b. Name each of the following compounds.  

(2.5 marks)

i. MnS\(_2\)  \[ \text{Manganese(IV)sulfide} \]

ii. HClO\(_4\) (aq)  \[ \text{perchloric acid} \]

iii. NH\(_4\)C\(_2\)H\(_3\)O\(_2\)  \[ \text{ammonium acetate} \]

iv. Li\(_3\)N  \[ \text{lithium nitride} \]

v. Na\(_2\)O\(_2\)  \[ \text{sodium peroxide} \]
Question 2  (6 marks)
Determine the empirical formula and the molecular formula of hydrogen tetrathionate (226.3 g/mol) that gives the following mass percentages upon analysis:

\[
\begin{align*}
\text{H} & = 0.91\% \\
\text{S} & = 56.67\% \\
\text{O} & = 42.42\%
\end{align*}
\]

For 100. g compound, we have:

\[
\begin{align*}
100 \text{ g} \times \frac{0.91}{100} \times 1 \text{ mol H} / 1.008 \text{ g} & = 0.903 \text{ mol H} \\
100 \text{ g} \times \frac{56.67}{100} \times 1 \text{ mol S} / 32.07 \text{ g} & = 1.767 \text{ mol S} \\
100 \text{ g} \times \frac{42.42}{100} \times 1 \text{ mol O} / 16.01 \text{ g} & = 2.650 \text{ mol O}
\end{align*}
\]

The molar stoichiometric ratio is:

\[
\begin{align*}
\text{H: } 0.903 / 0.903 & = 1.00 \text{ H} \\
\text{S: } 1.767 / 0.903 & = 1.96 \text{ S} \\
\text{O: } 2.650 / 0.903 & = 2.93 \text{ O}
\end{align*}
\]

Therefore, the empirical formula is: \( \text{HS}_2\text{O}_3 \).

Molar mass of the empirical formula = 113.1 g unit\(^{-1}\)

\[
\text{The ratio } = \frac{226.3 \text{ g.mol}^{-1}}{113.1 \text{ g.unit}^{-1}} = 2 \text{ units / mol}
\]

Therefore, the molecular formula is: \( 2(\text{HS}_2\text{O}_3) = \text{H}_2\text{S}_4\text{O}_6 \).

Ans. empirical formula: \( \text{HS}_2\text{O}_3 \) ..... Ans. molecular formula: \( \text{H}_2\text{S}_4\text{O}_6 \).
Question 3

Limestone, CaCO₃, reacts with hydrochloric acid to form calcium chloride, water and carbon dioxide according to the following reaction:

\[ \text{CaCO}_3(s) + 2\text{HCl}(aq) \rightarrow \text{CaCl}_2(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g) \]

How many liters of CO₂ gas will be formed at 755 torr and 33.0°C by the reaction of 2.35 g of limestone with an excess of hydrochloric acid? Assume 100% yield and that the gas is ideal.

*The yield is limited by the number of mole of CaCO₃.*

\[
2.35 \text{ g} \times \frac{1 \text{ mol}}{100.1 \text{ g}} = 2.348 \times 10^{-2} \text{ moles CaCO}_3 = 2.348 \times 10^{-2} \text{ moles CO}_2
\]

Since ideal gas: \[PV = nRT\]

*The volume of gas will be:*

\[
V = \frac{nRT}{P} \quad T = 273.2 + 33.0 = 306.2 \text{ K} \\
P = 755 \text{ torr} \times \frac{1.00 \text{ atm}}{760 \text{ torr}} = 0.993 \text{ atm}
\]

\[
V = (2.348 \times 10^{-2} \text{ mole})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(306.2 \text{ K}) / (0.993 \text{ atm})
\]

\[
V = 0.594 \text{ L of CO}_2 \text{ gas will be formed}
\]

Ans. volume of CO₂ gas formed: \[0.594 \text{ L}\]
Question 4

Cisplatin is an anticancer agent used for the treatment of solid tumors such as in breast cancer and is prepared by reacting potassium tetrachloroplatinate ($K_2PtCl_4$) and ammonia ($NH_3$). The other product generated is KCl.

Cisplatin

a. Write the balanced equation for this reaction

\[
K_2PtCl_4 + 2 NH_3 \rightarrow Pt(NH_3)_2Cl_2 + 2 KCl
\]

If 10.00 g of potassium tetrachloroplatinate is mixed with 0.500 g of ammonia ($NH_3$):

b. Identify the limiting reactant and calculate the theoretical yield.

\[
\text{Mole of Cisplatin produced if } K_2PtCl_4 \text{ is the limiting reactant:}
\]

\[
10.00 \text{ g } K_2PtCl_4 \times \frac{1 \text{ mol}}{415.10 \text{ g}} \times \frac{1 \text{ Cisplatin}}{1 \text{ K}_2\text{PtCl}_4} = 2.409 \times 10^{-2} \text{ mole Cisplatin}
\]

\[
\text{Mole of Cisplatin produced if ammonia is the limiting reactant:}
\]

\[
0.500 \text{ g } NH_3 \times \frac{1 \text{ mol}}{17.03 \text{ g}} \times \frac{1 \text{ Cisplatin}}{2 \text{ NH}_3} = 1.47 \times 10^{-2} \text{ mole Cisplatin}
\]

Therefore, $NH_3$ is the limiting reactant (least amount). The yield of Cisplatin is:

\[
1.47 \times 10^{-2} \text{ mole Cisplatin} \times \frac{300.1 \text{ g}}{\text{mole}} = 4.40 \text{ g Cisplatin}
\]

Ans. Limiting reactant: $NH_3$  Ans. theoretical yield: 4.40 g

c. If 3.52 g of Cisplatin is actually produced, what is the percent yield of Cisplatin?

\[
\frac{3.52 \text{ g}}{4.40 \text{ g}} \times 100\% = 79.9\%
\]

Ans. %yield of Cisplatin: 79.9%
Question 5

a. Complete and balance the following molecular equation. 

\[ \text{Fe(NO}_3\text{)}_2 \text{ (aq)} + 2 \text{NH}_4\text{OH (aq)} \rightarrow \text{Fe(OH)}_2(s) + 2\text{NH}_4\text{NO}_3\text{(aq)} \]

Give the physical state in brackets, ex: liquid = (l), for each of the substances participating to the reaction.
Write also the complete ionic equation (C.I.E.), identify the spectator ions and write the net ionic equation (N.I.E.) for this reaction.

C.I.E. \[ \text{Fe}^{2+}(aq) + 2\text{NO}_3^-(aq) + 2\text{NH}_4^+(aq) + 2\text{OH}^-(aq) \rightarrow \text{Fe(OH)}_2(s) + 2\text{NH}_4^+(aq) + 2\text{NO}_3^-(aq) \]

Spectator ions: \[ \text{NO}_3^-(aq) \text{ and } \text{NH}_4^+(aq) \]

N.I.E. \[ \text{Fe}^{2+}(aq) + 2\text{OH}^-(aq) \rightarrow \text{Fe(OH)}_2(s) \]

b. Classify each of the following reaction as: precipitation, acid-base or oxidation-reduction (red-ox)

<table>
<thead>
<tr>
<th>Chemical reaction</th>
<th>Classification</th>
</tr>
</thead>
<tbody>
<tr>
<td>i [ \text{C}_6\text{H}_12\text{O}_6(s) + 6\text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 6\text{H}_2\text{O(g)} ]</td>
<td>Red-ox</td>
</tr>
<tr>
<td>ii Lead(II) acetate(aq) + Sodium sulfate(aq) \rightarrow Lead(II) sulfate(s) + Sodium acetate(aq)</td>
<td>precipitation</td>
</tr>
<tr>
<td>iii Copper(II) Chloride(aq) + Zinc metal \rightarrow Cu(s) + ZnCl_2(aq)</td>
<td>Red-ox</td>
</tr>
</tbody>
</table>
Question 6

a. Assign oxidation state (or oxidation numbers) to nitrogen in each of the following compounds: (2 marks)

i. $\text{N}_2\text{H}_4$ ..........-2.......... iii. $\text{Na}_3\text{N}$ ..........-3.......... 

ii. $\text{N}_2\text{O}_3$ ..........+3.......... iv. $\text{NO}_3^-$ ..........+5.......... 

b. Balance the following oxidation-reduction reaction in an aqueous basic solution using the half-reaction method. (3 marks)

$$\text{ClO}_3^-(aq) + \text{N}_2\text{H}_4(aq) \rightarrow \text{NO}(g) + \text{Cl}^- (aq)$$

This equation has to be separated in two half-reactions first. Each half-equation is then balanced

$$\text{ClO}_3^-(aq) \rightarrow \text{Cl}^- (aq)$$

$$\ 6e^- + 6H^+ + \text{ClO}_3^-(aq) \rightarrow \text{Cl}^- (aq) + 3H_2O(l)$$

$$\text{N}_2\text{H}_4(aq) \rightarrow \text{NO}(g)$$

$$2H_2O(l) + \text{N}_2\text{H}_4(aq) \rightarrow 2\text{NO}(g) + 8H^+(aq) + 8e^-$$

Since there is no free electrons, the two half reactions are multiplied by a factor to cancel-out the electrons

$$(x4) \ \ 6e^- + 6H^+ + \text{ClO}_3^-(aq) \rightarrow \text{Cl}^- (aq) + 3H_2O(l)$$

$$(x3) \ \ 2H_2O(l) + \text{N}_2\text{H}_4(aq) \rightarrow 2\text{NO}(g) + 8H^+(aq) + 8e^-$$

$$24e^- + 24H^+ + 4\text{ClO}_3^-(aq) \rightarrow 4\text{Cl}^- (aq) + 12H_2O(l)$$

$$+ \ 6H_2O(l) + 3\text{N}_2\text{H}_4(aq) \rightarrow 6\text{NO}(g) + 24H^+(aq) + 24e^-$$

$$4\text{ClO}_3^-(aq) + 3\text{N}_2\text{H}_4(aq) \rightarrow 6\text{NO}(g) + 4\text{Cl}^- (aq) + 6H_2O(l)$$

Ans. balanced reaction: $4\text{ClO}_3^-(aq) + 3\text{N}_2\text{H}_4(aq) \rightarrow 6\text{NO}(g) + 4\text{Cl}^- (aq) + 6H_2O(l)$

c. Identify the substance being oxidized ..........$\text{N}_2\text{H}_4(aq)$........ (0.5 mark)

d. Identify the reducing agent ..........$\text{N}_2\text{H}_4(aq)$........ (0.5 mark)
**Question 7**

a. Calculate the de Broglie wavelength (in meters) of a proton that is traveling at 25.0% of the speed of light. 

\[
25.0\% \text{ of the speed of light is: } 25.0/100 \times 2.998 \times 10^8 \text{ m s}^{-1} = 7.50 \times 10^7 \text{ m s}^{-1}
\]

Since the de Broglie wavelength is given by \( \lambda = \frac{h}{mv} \), then

\[
\lambda = \frac{6.626 \times 10^{-34} \text{ J s}}{1.673 \times 10^{-27} \text{ kg} \times 7.50 \times 10^{-7} \text{ m s}^{-1}} = 5.28 \times 10^{-15} \text{ m}
\]

Ans. de Broglie wavelength: \( 5.28 \times 10^{-15} \text{ m} \).

b. One of the emission lines of the hydrogen atom has a wavelength of 656.3 nm.

i. In what region of the electromagnetic spectrum (ex: X-ray, UV, etc.) is this emission line found?

Ans: Emission at 656.3 nm = visible

ii. Calculate the energy of the photon emitted by this transition.

\[
E = \frac{hc}{\lambda}
\]

Therefore, the energy of the photon is:

\[
E = 6.626 \times 10^{-34} \text{ J s} \times \frac{2.998 \times 10^8 \text{ m s}^{-1}}{656.3 \times 10^{-9} \text{ m}} = 3.027 \times 10^{-19} \text{ J}
\]

iii. Determine the initial or the final \( n \) value associated with this emission line if one of the two energy levels is \( n = 2 \).

Since this problem is about the hydrogen atom, the Balmer-Rydberg equation can be used. However, the Bohr’s model of the atom is also appropriate here:

\[
E = -\frac{BZ^2}{n^2} \quad \text{Since } Z = 1 \text{ for the hydrogen atom and } |\Delta E_{\text{atom}}| = E_{\text{photon}}, \text{ then}
\]

\[
E_{\text{photon}} = \left| -B \left( \frac{1}{n_{\text{final}}^2} - \frac{1}{n_{\text{initial}}^2} \right) \right|
\]

and \( 3.027 \times 10^{-19} \text{ J} = 2.18 \times 10^{-18} \text{ J} \left( \frac{1}{2^2} - \frac{1}{n_{\text{initial}}^2} \right) \text{ then } n_{\text{initial}} = 3 \)
Question 8

a. Supply the missing quantum number(s) or sublevel names (or subshell):

<table>
<thead>
<tr>
<th></th>
<th>n</th>
<th>l</th>
<th>m_l</th>
<th>subshell</th>
</tr>
</thead>
<tbody>
<tr>
<td>i</td>
<td>4</td>
<td>1</td>
<td>0</td>
<td>4p</td>
</tr>
<tr>
<td>ii</td>
<td>3</td>
<td>2</td>
<td>-2</td>
<td>3d</td>
</tr>
<tr>
<td>iii</td>
<td>2</td>
<td>0</td>
<td>0</td>
<td>2s</td>
</tr>
</tbody>
</table>

b. Consider the element with the ground-state electron configuration $1s^22s^22p^4$

i. Identify this element (name or atomic symbol) (0.5 mark)

Oxygen or “O”

ii. Write its electron configuration using the orbital diagram (box notation) (0.5 mark)

\[ \begin{array}{c}
\uparrow \downarrow \\
1s & 2s & 2p \\
\end{array} \]

c. Consider the atom chromium, Cr

i. Write its complete ground-state electron configuration (0.5 mark)

\[ \begin{array}{c}
\uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \\
1s & 2s & 2p & 3s & 3p & 4s & 3d & 4p \\
\end{array} \]

Therefore: $1s^22s^22p^63s^23p^64s^13d^5$

ii. What is the number of unpaired electrons in Cr atom? (0.5 mark)

6 unpaired electrons in Chromium

d. Give the set of quantum numbers for the circled electron in the following orbital diagram: (1 mark)

\[ \begin{array}{c}
\uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \uparrow \downarrow \\
1s & 2s & 2p & 3s & 3p \\
\end{array} \]

$n = 3$, $l = 1$, $m_l = -1, 0, +1$, $m_s = \frac{1}{2}$
Question 9

Indicate on this blank outline of a periodic table, which element is described by each of the following. Place the letter of the question (a, b, c, etc) in the correct element box.

*Messy answers will not be marked.*

a. Smallest atomic radius in Group 6A
b. Largest atomic radius in Period 6
c. Condensed ground-state electron configuration is [Ne] 3s²3p²
d. The period 4 member whose (2-) ion is isoelectronic with Kr
e. A transition metal ion with a charge of 1+ having 5 unpaired “4d” electrons
f. The element with the highest first ionization energy in period 4
g. The excited electron configuration is 1s²2s²2p⁶3s¹3p¹
h. The halogen with the smallest electron affinity (less exothermic)
i. The noble gas with electrons occupying 4f orbitals
j. The least electronegative transition metal in period 5
Question 10

a. Draw all the possible Lewis structures (resonance forms) for the following ion with carbon as the central atom: \( \text{SCN}^- \) (3 marks)

b. Assign formal charges to each atom in each structure. (2 marks)

c. Indicate the most appropriate structure according to the formal charges? (1 mark)

The most stable one

\[ \begin{align*}
: & \equiv C - \overset{-2}{N} : \\
\text{(+1)} & \quad \text{(0)} & \quad \text{(-2)} & \quad \text{(0)} & \quad \text{(-1)} & \quad \text{(0)} & \quad \text{(0)} \\
\overset{0}{S} & \equiv C & \overset{0}{N} & \quad \text{(-1)} & \quad \text{(0)} & \quad \text{(0)} & \quad \text{(0)}
\end{align*} \]

The most appropriate structure
since the formal charge (-1)
is on the most electronegative atom
**Question 11**

(9 marks)

Complete the table below. For each of the following molecules:

a. Draw the Lewis structure. Also give the name of the electron-pairs arrangement.

b. Make a “3D”-sketch of the molecule with the bond angles and provide a name for the molecular structure (or shape of the molecule).

<table>
<thead>
<tr>
<th>Molecule</th>
<th>Lewis structure with all the lone pairs name of the electron-pairs arrangement</th>
<th>3D-sketch (VSEPR) with bond angles name of the shape of the molecule</th>
<th>Is this a polar molecule? (Y / N)</th>
</tr>
</thead>
<tbody>
<tr>
<td>PH₃</td>
<td>$5+3\times 1 = 8$ valence electrons</td>
<td><img src="image" alt="3D-sket" /></td>
<td>No since $EN = 2.1$ for both H and P</td>
</tr>
<tr>
<td>BrF₃</td>
<td>$4\times 7 = 28$ valence electrons</td>
<td><img src="image" alt="3D-sket" /></td>
<td>yes</td>
</tr>
<tr>
<td>SO₂</td>
<td>$6\times 3 = 18$ valence electrons</td>
<td><img src="image" alt="3D-sket" /></td>
<td>yes</td>
</tr>
</tbody>
</table>
**Question 12**
The structure for the amino acid methionine is given below.

a. Complete the structure by adding all missing lone pairs of electrons.  

b. Complete the table for each of the atoms labeled 1 to 6. For each labeled atom, give the hybridization, the bond angle and the number of σ and π bonds around the atom.

<table>
<thead>
<tr>
<th>Atom 1</th>
<th>Atom 2</th>
<th>Atom 3</th>
<th>Atom 4</th>
<th>Atom 5</th>
<th>Atom 6</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hybridization</td>
<td>$sp^3$</td>
<td>$sp^3$</td>
<td>$sp^2$</td>
<td>$sp^2$</td>
<td>$sp^3$</td>
</tr>
<tr>
<td>Bond angle</td>
<td>109.5</td>
<td>$&lt;109.5$</td>
<td>120</td>
<td>no angle</td>
<td>$&lt;109.5$</td>
</tr>
<tr>
<td>Number of σ bonds around the atom</td>
<td>4</td>
<td>2</td>
<td>3</td>
<td>1</td>
<td>2</td>
</tr>
<tr>
<td>Number of π bonds around the atom</td>
<td>0</td>
<td>0</td>
<td>1</td>
<td>1</td>
<td>0</td>
</tr>
</tbody>
</table>

**Question 13**

a. Write the name of the most important intermolecular force that keeps the following molecules in the liquid phase.  

i. CH$_3$F  ................................................. dipole-dipole  .................................................

ii. H$_2$O  ................................................. H-bond  .................................................

iii. Br$_2$  ................................................. London dispersion forces  .................................................

b. Arrange the following three substances in order of increasing boiling point.  

HCl, HF, HBr

\[
\begin{align*}
\text{HCl} & \quad \text{Low boiling point} \\
\text{HBr} & \quad \text{Low boiling point} \\
\text{HF} & \quad \text{High boiling point}
\end{align*}
\]
**Question 14**

a. Sketch the phase diagram for carbon dioxide (CO$_2$) from the following data: (4 marks)
   
i. The triple point is at 5.2 atm, and −57°C  
ii. The critical point is at 72.8 atm, and 31°C.  
iii. At a pressure of 1 atm, the solid-gas phase transition takes place at −78°C  
iv. At a pressure of 72.8 atm, the solid-liquid phase transition occurs at −21°C

*Don’t forget to label your axes and to indicate the phase (ex: g, l, s) in each region*

![Phase Diagram of CO$_2$](image)

From your phase diagram, answer the following questions:

b. For CO$_2$ at 5 atm and −50°C, what is the stable phase present (gas, liquid, etc.)? (1 mark)

   *Gas phase (check graph)*

c. What phase changes occur when the pressure of a sample of CO$_2$ is decreased from 70 atm to 7 atm at a constant temperature of 0°C? (1 mark)

   *Liquid phase to gas phase*
**Question 15**

a. Calculate $\Delta H$ for the overall reaction:  

$$4\text{NO}_2(g) + \text{O}_2(g) \rightarrow 2\text{N}_2\text{O}_5(s)$$

using the following data

equation 1: $2\text{N}_2\text{O}_5(s) \rightarrow 4\text{NO}(g) + 3\text{O}_2(g) \quad \Delta H = 447.4 \text{ kJ}$
equation 2: $2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g) \quad \Delta H = -114.2 \text{ kJ}$

Need to reverse equation 2 and double the amount:

equation 3: so: $4\text{NO}_2(g) \rightarrow 4\text{NO}(g) + 2\text{O}_2(g) \quad \Delta H = +228.4 \text{ kJ}$

Need to reverse the equation 1:

Equation 4: $4\text{NO}(g) + 3\text{O}_2(g) \rightarrow 2\text{N}_2\text{O}_5(s) \quad \Delta H = -447.4 \text{ kJ}$

Solution: need to add the equation 3 and 4 and the corresponding $\Delta H$:

answer: $4\text{NO}_2(g) + \text{O}_2(g) \rightarrow 2\text{N}_2\text{O}_5(s) \quad \Delta H = -219.0 \text{ kJ}$

Ans. $\Delta H$ overall reaction: ........-219.0 kJ........

b. For the reaction below:

$$\text{H} \equiv \text{C} \equiv \text{N}(g) + 2\text{H}_2(g) \rightarrow \text{H} - \text{C} - \text{N} \equiv \text{H}(g) \quad \Delta H = -158 \text{ kJ}$$

Calculate the bond energy for the $\text{C} \equiv \text{N}$ triple bond from the following data:  

<table>
<thead>
<tr>
<th>Bond</th>
<th>Bond energy (kJ.mol$^{-1}$)</th>
<th>Bonds broken</th>
<th>Bonds formed</th>
</tr>
</thead>
<tbody>
<tr>
<td>H–H</td>
<td>432</td>
<td>$C-H = 413 \text{ kJ/mol}$</td>
<td>$C-N = 305 \text{ kJ/mol}$</td>
</tr>
<tr>
<td>C–H</td>
<td>413</td>
<td>$C \equiv H ?$</td>
<td>$3 C-H = 3x (-413 \text{ kJ/mol})$</td>
</tr>
<tr>
<td>C–N</td>
<td>305</td>
<td></td>
<td></td>
</tr>
<tr>
<td>N–H</td>
<td>391</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Since $\Delta H = -158 \text{ kJ}$ then:

$$-158 \text{ kJ} = 1277 \text{ kJ} + C \equiv H - 2326 \text{ kJ}$$

Ans. $\text{C} \equiv \text{N}$ bond energy: 891 kJ/mole
**Question 16**

You want to determine the Cu(II) concentration in an unknown solution by titration. You will first precipitate Cu$^{2+}$ from the solution by adding an excess of iodide ion

$$2 \text{Cu}^{2+}(aq) + 4 I^{-}(aq) \rightarrow 2 \text{CuI(s)} + I_2(aq) \quad \text{[unbalanced]}$$

The corresponding amount of iodine (I$_2$) formed will then be titrated with a sodium thiosulfate solution:

$$2 \text{S}_2\text{O}_3^{2-}(aq) + I_2(aq) \rightarrow S_4\text{O}_6^{2-}(aq) + 2 I^{-}(aq) \quad \text{[unbalanced]}$$

The end-point of the titration is reached when all the I$_2$ has reacted. At this point, the violet starch indicator has turned white. Use the following values from the laboratory data sheet to calculate the concentration of Cu$^{2+}$ in the unknown.

**Volumetric analysis of Cu$^{2+}$**

**DATA SHEET**

a. Write the overall stoichiometric equation

$$2 \text{Cu}^{2+}(aq) + 2 I_2(aq) + 2 \text{S}_2\text{O}_3^{2-}(aq) \rightarrow 2 \text{CuI(s)} + S_4\text{O}_6^{2-}(aq)$$

b. Results and observations

Unknown no: 20

Concentration of the standard Na$_2$S$_2$O$_3$ solution, M $\text{0.0206}$

Initial burette reading, mL $\text{0.17}$

Final burette reading (mL) $\text{9.82}$

Volume of Na$_2$S$_2$O$_3$ used (mL) $\text{9.65}$

Number of mole of Na$_2$S$_2$O$_3$ used $\text{1.99x10}^{-4}$

Number of mole of Cu$^{2+}$ in the unknown sample $\text{1.99x10}^{-4}$

Volume of the original Cu$^{2+}$ solution used (mL) $\text{10.00}$ (from a 10 mL pipette)

Concentration of the Cu$^{2+}$ in the original solution $\text{(M)} \ 1.99x10^{-2}$

**Concentration of Cu$^{2+}$ in the original solution**

$$\frac{1.99x10^{-4} \text{ mole Cu}^{2+}}{10.00 \text{ mL}} \times \frac{1000 \text{ mL}}{L} = 1.99x10^{-2} \text{ M Cu}^{2+}$$