

## Cambridge Chemistry Challenge Lower 6th

June 2021

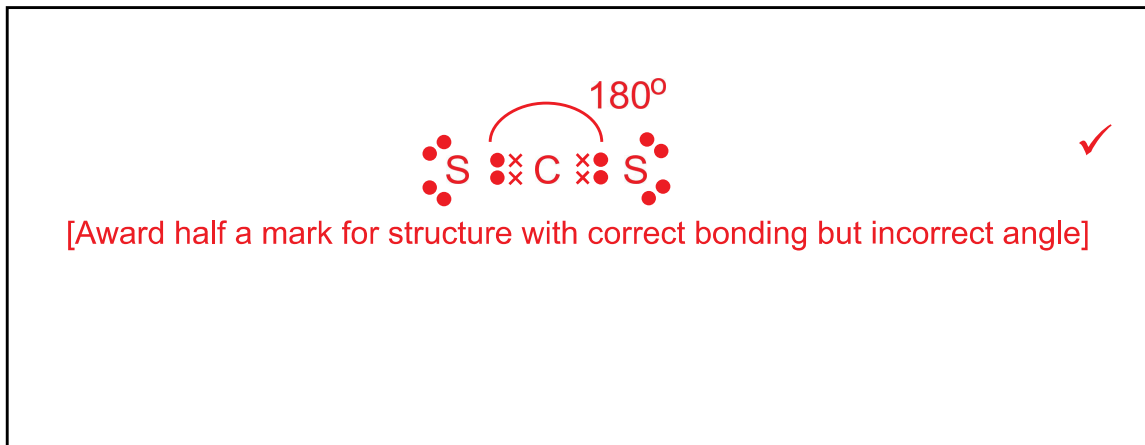
# Marking scheme for teachers

(please also read the additional instructions)

	p2	p3	p4	p5	p6	p7	p8	<b>Total</b>
mark	6	8	10	8	13	10	5	60

1(a) (i) Dot-and-cross diagram of carbon disulfide & bond angle:

1



(ii) Equation for the reaction between ammonia and carbon disulfide:

1



[Do not penalise lack of state symbols throughout.  
Could give formula for  $\text{NH}_4\text{SCN}$  as  $\text{CH}_4\text{N}_2\text{S}$ .]

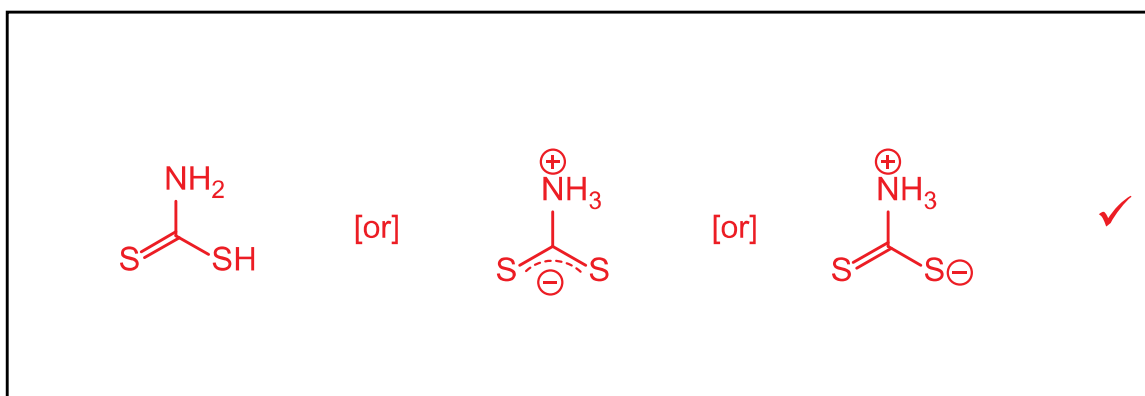
(iii) Molecular formula for intermediate compound **A**:

1



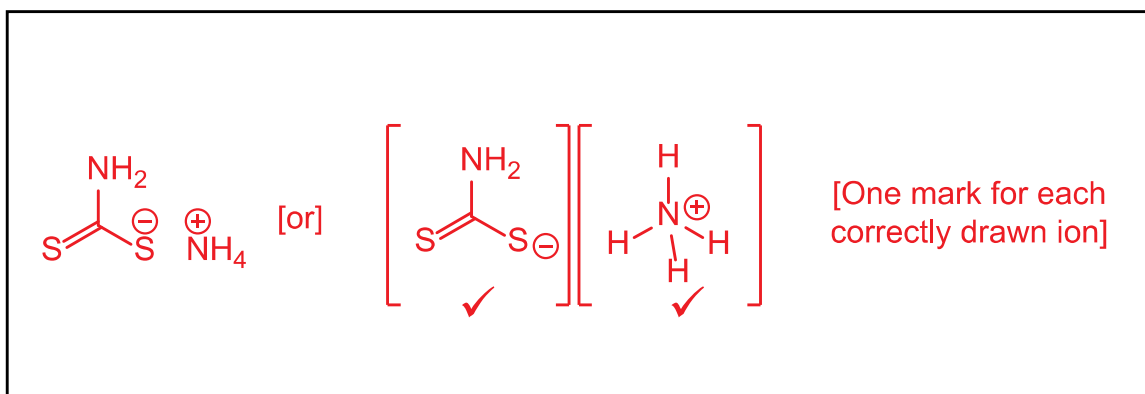
1(b) (i) Structure of reactive species **X**:

1

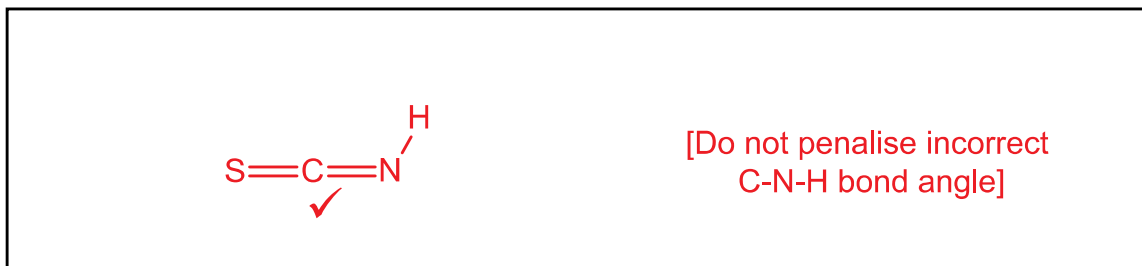


(ii) Structure of intermediate compound **A**:

2



1(c) (i) Structure of isothiocyanic acid:



1

(ii) Structure of thiocyanic acid:



1

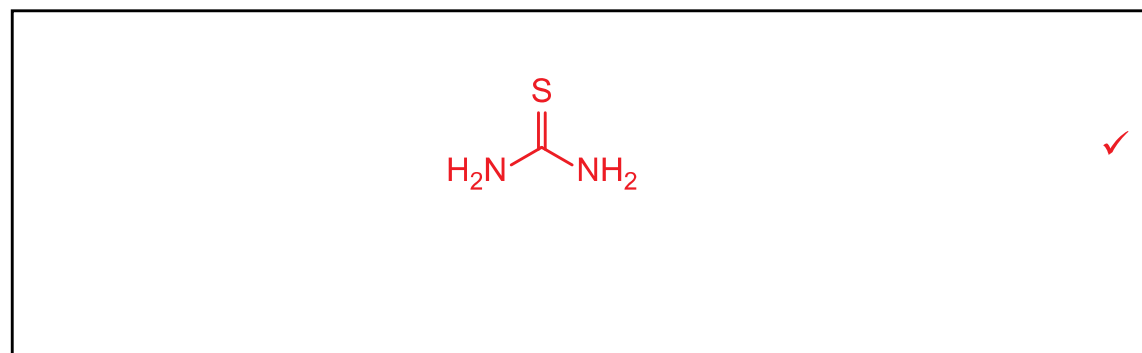
(iii) Bonds in order of INCREASING length



[No half marks, no ECF from part (i) or (ii)]

1

1(d) Structure of compound **C**:



1

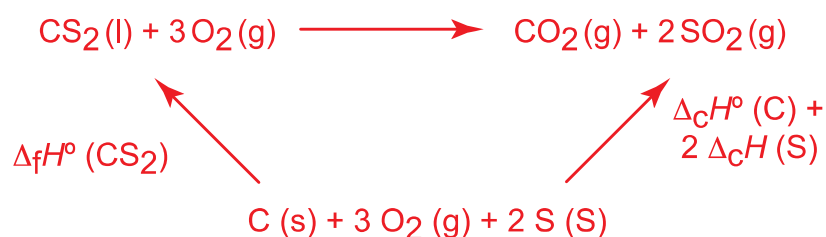
1(e) (i) Equation for combustion of 1 mole of CS<sub>2</sub>:



1

(ii) Standard enthalpy of combustion for CS<sub>2</sub>:

2



$$\begin{aligned}
 \Delta_c H^\circ &= \Delta_c H^\circ(C) + 2\Delta_c H^\circ(S) - \Delta_f H^\circ(CS_2) \quad \checkmark \\
 &= (-393.5) + (2 \times -297.5) - (89.4) \\
 &= -1080 \text{ kJ mol}^{-1} \quad (3 \text{ sig. fig.}) \quad \checkmark
 \end{aligned}$$

(iii) Equation for the decomposition of mercury(II) thiocyanate:



1

1(f) Percentage mass loss for mercury(II) sulfide and carbon nitride and identity of compound removed between 245 and 370°C:

5

[Note CS<sub>2</sub> does not appear in the serpent ash as it burns on formation]

$$M_r \text{ HgS} = 200.59 + 32.06 = 232.65$$

$$M_r \text{ C}_3\text{N}_4 = (3 \times 12.01) + (4 \times 14.01) = 92.07$$

$$\text{"M}_r \text{ Serpent"} = (2 \times M_r \text{ HgS}) + M_r \text{ C}_3\text{N}_4 = 557.37$$

$$\% \text{ Mass loss of HgS} = (2 \times 232.65 / 557.37) \times 100 = 83.5\% \quad \checkmark\checkmark$$

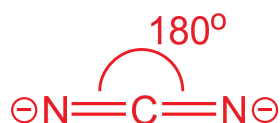
$$\% \text{ Mass loss of C}_3\text{N}_4 = (92.07/557.37) \times 100 = 16.5\% \quad \checkmark\checkmark$$

Mercury(II) sulfide must be lost in this temperature range.  $\checkmark$

[If forgot that CS<sub>2</sub> has burned off, calculate loss of HgS as 73.4% and C<sub>3</sub>N<sub>4</sub> as 14.5%. Give 1 mark each for these answers.]

1(g) (i) Structure of CN<sub>2</sub><sup>2-</sup> & bond angle.

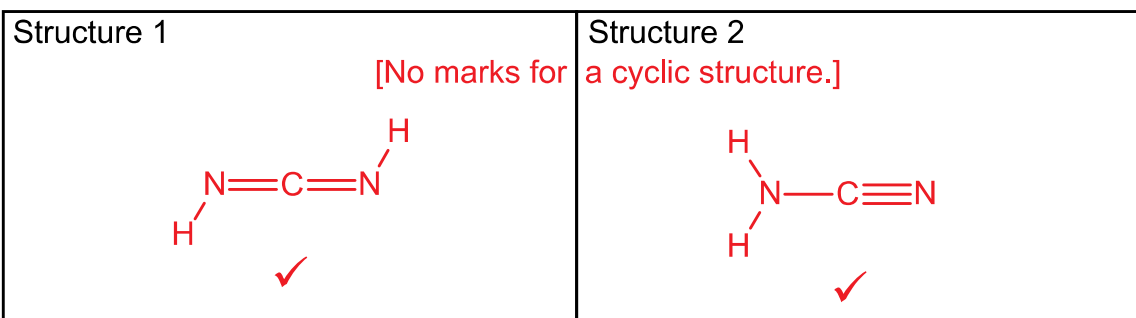
1



[Award half a mark for structures with correct bonding but incorrect angle]

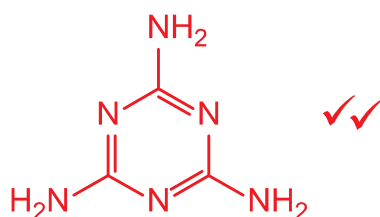
(ii) Two possible Structures for cyanamide:

2



(iii) Structure of melamine:

2



1(h) (i) RMM of the molecule lost when melamine dimerises:

2

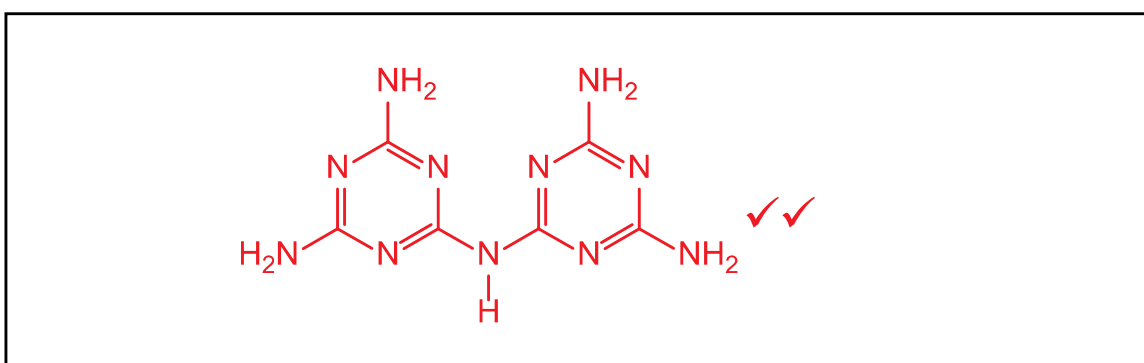
$M_r \text{ melamine} = 3 \times A_r \text{ C} + 6 \times A_r \text{ H} + 6 \times A_r \text{ N}$   
 $= (3 \times 12.01) + (6 \times 1.008) + (6 \times 14.01) = 126.138$   
 If no atoms were lost the  $M_r$  of the melamine dimer =  $2 \times 126.138 = 252.276$  ✓  
 6.75% of mass is  $0.0675 \times 252.276 = 17.0$  ✓  
 (3 sig figs)

(ii) Molecule **M** lost when melamine dimerises: ammonia,  $\text{NH}_3$ . ✓

1

(iii) Structure of melamine dimer:

2



1(j) Tick one box that best explains the presence of hydrogen in **g-CN**:

1

- Hydrogen gas is easily absorbed from the atmosphere inbetween layers of **g-CN**.
- Polymerisation reactions inevitably leave some hydrogen attached to N-atoms at the edges of the sheets of **g-CN**.
- g-CN** is very reactive and reacts with atmospheric moisture to attach hydrogen atoms.
- There are contaminants in the original melamine that contain hydrogen and these react with the **g-CN**.

1(k) Empirical formula of the carbon nitride ultimately formed on heating melamine:

2

molecular formula of melamine is  $\text{C}_3\text{H}_6\text{N}_6$   
 If all H atoms are lost as ammonia, need to lose 6H with 2N  
 so empirical formula of the carbon nitride is  $\text{C}_3\text{N}_4$  (as before). ✓✓

2(a) (i) Percentage decrease in mass when copper(II) oxide is reduced to copper:

$$\begin{aligned} \text{Percentage mass loss} &= (A_r \text{ O} / M_r \text{ CuO}) \times 100 \\ &= 16.00 / (63.55 + 16.00) \times 100 \\ &= 20.1 \% \text{ (3 sig. fig.) } \checkmark \end{aligned}$$

1

(ii) Two equations for the production of copper from copper(II) oxide and carbon:



[Do not penalise lack of state symbols throughout]

2

(iii) Minimum theoretical mass of carbon needed to produce 100g copper:

$$\begin{aligned} \text{Number of moles of Cu produced} &= \text{mass Cu} / A_r \text{ Cu} \\ &= 100 / 63.55 \\ &= 1.574 \text{ moles. } \checkmark \end{aligned}$$

When CO is produced one mole of C produces one mole of Cu.  
When CO<sub>2</sub> is produced one mole of C produces two moles of Cu.

Assuming CO<sub>2</sub> production, 1.574 / 2 = 0.7868 moles of C are needed.  
Mass of C = moles C x A<sub>r</sub> C = 0.7868 x 12.01 = 9.45g  $\checkmark$

2

2(b) Oxide formulae:



[Award half a mark per correct formula]

3

2(c) (i) Molecular or ionic formulae:



[Award half a mark per correct formula]

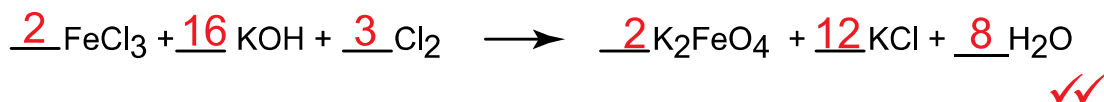
(ii) Formula of the non-tetrahedral molecule or ion:



1

2(d) (i) Balanced equation for the preparation of potassium ferrate:

2



[Allow fractions if still balanced correctly.]

(ii) Balanced equation for the decomposition of aqueous potassium ferrate:

2



[Allow fractions if still balanced correctly.]

(iii) Order of pH 4, pH 7 and pH 10 from slowest to fastest expected reaction:

1

pH 10                      pH 7                      pH 4                      ✓  
 slowest                                           fastest

2(e) (i) Mass of manganese(II) sulfate monohydrate in 750 ml sample:

3

$$\begin{aligned} \text{Number of moles of Mn(II) in 750 ml} &= [\text{Mn(II)}] \times \text{volume} \\ &= 9.6 \times 10^{-6} \times (750/1000) \\ &= 7.2 \times 10^{-6} \text{ moles } \checkmark \end{aligned}$$

$$\begin{aligned} M_r \text{ MnSO}_4 \cdot \text{H}_2\text{O} &= A_r \text{ Mn} + A_r \text{ S} + (5 \times A_r \text{ O}) + (2 \times A_r \text{ H}) \\ &= 54.94 + 32.06 + (5 \times 16.00) + (2 \times 1.008) \\ &= 169.016 \checkmark \end{aligned}$$

$$\begin{aligned} \text{Mass MnSO}_4 \cdot \text{H}_2\text{O} &= \text{Number of moles of Mn(II)} \times M_r \text{ MnSO}_4 \cdot \text{H}_2\text{O} \\ &= 7.2 \times 10^{-6} \times 169.016 \\ &= 0.00122 \times 10^{-3} \text{ g} \\ &= 1.2 \text{ mg (2 sig. fig.) } \checkmark \end{aligned}$$

[Award 2 marks for 1220 µg or 1.2 x 10<sup>-3</sup> g]  
 [Award 1 mark for 1.2 g]

(ii) Concentration of potassium ferrate stock solution:

2

$$\begin{aligned} \text{Stock solution: } [\text{K}_2\text{FeO}_4]_a; v_a &= 0.0080 \text{ dm}^3 \\ \text{Stock + Mn(II) solution: } [\text{K}_2\text{FeO}_4]_b &= 9.6 \times 10^{-6} \text{ mol dm}^{-3}; v_b = 0.758 \text{ dm}^3 \\ [\text{K}_2\text{FeO}_4]_a &= ([\text{K}_2\text{FeO}_4]_b \times v_b) / v_a \checkmark \\ &= (9.6 \times 10^{-6} \times 0.758) / 0.008 \\ &= 9.096 \times 10^{-4} \text{ mol dm}^{-3} \text{ or } 910 \text{ } \mu\text{M } \checkmark \end{aligned}$$

2(f) (i) Molar ratio ferrate ions : Mn(II) ions:

1

The gradient of the graph is 0.75

For every 1  $\mu\text{M}$  increase in initial  $[\text{K}_2\text{FeO}_4]$  the  $[\text{Mn(II)}]$  lost increases by 1.5  $\mu\text{M}$

Molar ratio ferrate ions : Mn(II) ions = 2 : 3 ✓

(ii) Number of electrons formally transferred during the reaction:

1

Six electrons are formally transferred as per half-equation below ✓



(iii) Oxidation state for manganese in the product:

1

The product contains Mn(IV) as per the half equation below: ✓



(iv) Equation for reaction between potassium ferrate and manganese(II) chloride:

2

