1. The boiling point of hydrogen bromide is –67 ºC.
The boiling point of hydrogen iodide is –34 ºC.

The different boiling points can be explained in terms of the strength of bonds or interactions.

Which bonds or interactions are responsible for the higher boiling point of hydrogen iodide?

A covalent bonds

B hydrogen bonds

C permanent dipole–dipole interactions

D induced dipole–dipole interactions

Your answer D

[1]

2. This question is about ions and compounds containing hydrogen.

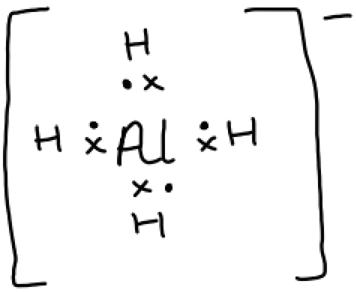
(a) Lithium aluminium hydride, $LiAlH_4$, contains the AlH_4^- ion.

ionic bonding

Draw a '*dot-and-cross*' diagram to show the bonding in an AlH_4^- ion.

covalent bonding in the ion

Show outer electrons only.



[1]

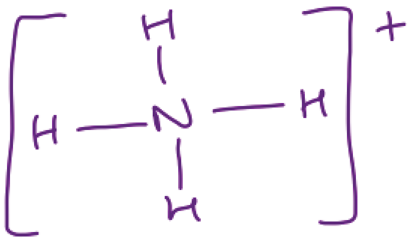
(b) Nitrogen forms NH_4^+ and NH_2^- ions.

Predict the name of the shape of, and H–N–H bond angle in, NH_4^+ and NH_2^-.

Ion	Name of shape	H–N–H bond angle
NH_4^+	tetrahedral	109.5°
NH_2^-	non-linear	104.5°

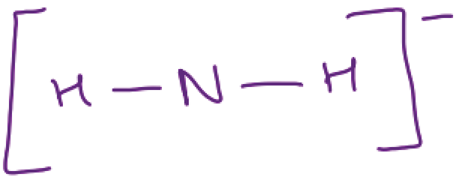
[2]

NH_4^+



4bp
0lp

NH_2^-



2bp
2lp

(c) Nitrogen, phosphorus and arsenic are in Group 15 (5) of the periodic table.

The boiling points of their hydrides are shown below.

hydrogen bonding between:
N – H
O – H
F – H

Element	Hydride	Boiling point / °C
N	NH_3	–33
P	PH_3	–88
As	AsH_3	–55

(i) Explain why the boiling point of PH_3 is lower than the boiling point of NH_3.

NH_3 has hydrogen bonding and PH_3 doesn't have hydrogen bonding. More energy is required to overcome hydrogen bonds so NH_3 has a higher boiling point.

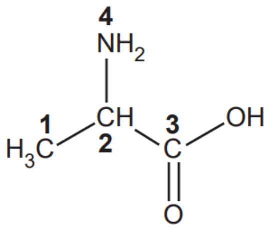
[2]

(ii) Explain why the boiling point of PH_3 is lower than the boiling point of AsH_3.

AsH_3 has more electrons than PH_3 and AsH_3 has stronger induced dipole - dipole interactions.

[2]

3. Four atoms, **1–4**, are labelled in the structure below.



Which atom has a trigonal planar arrangement of bonds around it?

→ 3 bonded pairs (bp) with no lone pairs (lp)

A Atom 1 4bp 0lp

B Atom 2 4bp 0lp

C Atom 3 3bp 0lp

D Atom 4 3bp 1lp

Your answer C

[1]

4. Within the permafrost in Arctic regions of the Earth, large amounts of methane are trapped within ice as 'methane hydrate', $CH_4 \cdot xH_2O$. Methane makes up about 13.4% of the mass of 'methane hydrate'.

Scientists are concerned that global warming will melt the permafrost, releasing large quantities of methane into the atmosphere.

tetrahedral
4 bonded pairs
non-linear
2 bonded pairs
2 lone pairs

(a) The H–O–H bond angle in ice is about 109° but about 105° in gaseous H_2O.

Explain why there is this difference.

Ice can form more hydrogen bonds. Water (H_2O) has 2 bonded pairs and 2 lone pairs:

lone pairs repel more than bonded pairs hence different bond angles. **[3]**

(b) Why are scientists concerned about the release of methane into the atmosphere?

Contributes towards global warming **[1]**

(c) Determine the formula of 'methane hydrate', $CH_4 \cdot xH_2O$.

In the formula, show the value of x to **two** decimal places.

$$\frac{13.4}{(12+4)} = \frac{100}{(12+4)+18x}$$

$$13.4(16+18x) = 100(16)$$

$$214.4 + 241.2x = 1600$$

$$241.2x = 1385.6$$

$$x = 5.74$$

formula = $CH_4 \cdot 5.74H_2O$ **[2]**

(d) Calculate the volume of methane, in dm^3, that would be released from the melting of each 1.00 kg of 'methane hydrate' at 101 kPa and 0 °C.

Give your answer to **three** significant figures.

$PV = nRT$ (P: Pa, V: m^3, T: K)

moles of $CH_4 \cdot 5.74H_2O$

$$\frac{1 \times 10^3}{12+4+(5.74 \times (2+16))} = 8.38 \text{ mol}$$

$$V = \frac{nRT}{P}$$

$$V = \frac{8.38 \times 8.314 \times 273}{101 \times 10^3} = 0.188 \text{ m}^3 = 188 \text{ dm}^3$$

volume = 188 dm^3 **[4]**

(e) Suggest why some industries are interested in the presence of 'methane hydrate' in regions of the Earth.

for fuel / energy

[1]

5. In the Periodic Table, element **X** is in Group 2 and element **Y** is in Group 15 (5).

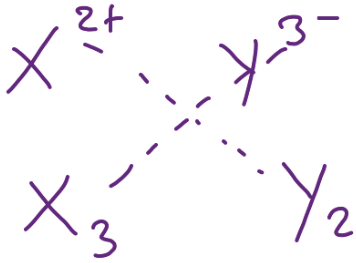
What is the likely formula of an ionic compound of **X** and **Y**?

A X_2Y_5

B X_2Y_3

C X_3Y_2

D X_5Y_2



answer C **[1]**

6. This question refers to the elements in the first three periods (H→Ar) of the Periodic Table.

(a) Select an element from the first three periods that fits each of the following descriptions.

(i) The element that forms a 1– ion with the same electron configuration as helium.

H gains an electron **[1]**

(ii) The element with the highest first ionisation energy.

He ionisation energy increases from the bottom left to top right of the periodic table **[1]**

(iii) The element in Period 3 which has the successive ionisation energies shown below.

Ionisation number	1st	2nd	3rd	4th
Ionisation energy/kJ mol^{-1}	738	1451	7733	10541

period 3 group 2 large jump meaning 2+ charge

Mg **[1]**

(iv) The element which forms a compound with fluorine that has octahedral molecules.

S **[1]**

(v) An element which reacts with water to form an acidic solution.

Cl or F forms HCl, HOCl HF, HOF **[1]**

(vi) The element **X**, which forms a compound with hydrogen, XH_3, with a molar mass of 34.0 g mol^{-1}.

$x = 34 - 3 = 31$ (RFM of x)

P **[1]**

(vii) An element which forms a compound with hydrogen in which the element has an oxidation number of –4.

group 4

C **[1]**

(viii) The element which has a density of 1.33×10^{-3} g cm^{-3} at room temperature and pressure.

O_2 $1.33 \times 10^{-3} \times 24000 \ cm^3$ (volume at RTP) **[1]**

$= 32$ (RFM of element)

(b) **Table 1.1** shows some properties of Period 3 chlorides.

Group		1	2	14 (4)	15 (5)	16 (6)
Chloride		$NaCl$	$MgCl_2$	$SiCl_4$	PCl_3	SCl_2
Electrical conductivity	**Solid**	poor	poor	poor	poor	poor
	Liquid	good	good	poor	poor	poor
Melting point		high	high	low	low	low

Table 1.1

Explain the properties shown in **Table 1.1** in terms of bonding and structure.

NaCl + MgCl2:
giant ionic lattice, ions are mobile in liquid state so are better electrical conductors

SiCl4, PCl3, SCl2:
simple molecular so held together by london forces of attraction

ionic bonds are stronger than london forces hence NaCl and MgCl2 have higher melting points. **[5]**

7. Which set of elements in the solid state contain a simple molecular lattice, a giant covalent lattice and a giant metallic lattice?

A S, Si, A*l*

B P, Si, C

C S, P, Si

D Mg, P, S

Your answer A **[1]**

8. These short questions are from different areas of chemistry.

(a) Explain why a CF_4 molecule has polar bonds but does **not** have an overall dipole.

F is more electronegative than C but CF_4 is symmetrical so the dipoles cancel out **[2]**

(b) Explain why a small proportion of molecules in water have a relative molecular mass of 20.

molecules contain: ^2H (deuterium)
^3H (tritium)

[1]

(c) What is the partial pressure of O_2 (in Pa) in a gas mixture containing 21% O_2 by volume and with a total pressure of 1.0×10^5 Pa?

partial pressure = mol fraction × total pressure

$0.21 \times 1.00 \times 10^5 = 2.1 \times 10^4$ Pa

21% = 0.21 = mol fraction

partial pressure of O_2 = 2.1×10^4 Pa **[1]**

(d) What mass of carbon dioxide (in g) is formed by the complete combustion of 42.0 m^3 (measured at RTP) of propane?

mol = vol (dm^3) / molar volume = 24 dm^3 at RTP

$\frac{42 \times 10^3}{24} = 1750$ mol of C_3H_8

$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

$(1750 \times 3) \times (12 + (16 \times 2)) = 2.31 \times 10^5$ g

mass = 2.31×10^5 g **[2]**

(e) A reaction is first order with respect to H^+. At a pH of 1, the initial rate is 2.4×10^{-3} mol dm^{-3} s^{-1}.

What is the initial rate at a pH of 3?

$\frac{2.4 \times 10^{-3}}{10^2} = 2.4 \times 10^{-5}$ mol dm^{-3} s^{-1}

initial rate = 2.4×10^{-5} mol dm^{-3} s^{-1} **[1]**

mass

mol × RFM

(f) What is the number of oxygen atoms in 4.26 g of P_2O_5?

$$\frac{4.26}{(31\times2)+(16\times5)} = 0.03 \text{ mol of } P_2O_5$$

$$(5\times0.03)\times6.02\times10^{23} = 9.03\times10^{22}$$

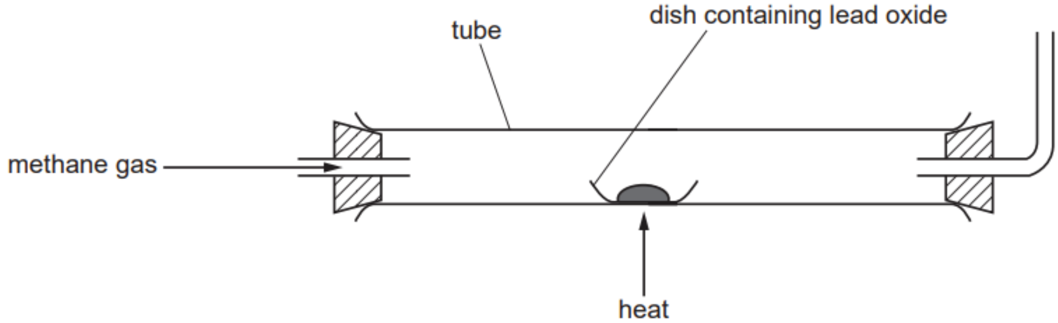
number of oxygen atoms = 9.03×10^{22} **[2]**

9. This question is about elements and compounds in Group 14 (Group 4) of the periodic table.

(a) There are four oxides of lead: PbO, PbO_2, Pb_2O_3 and Pb_3O_4.

A student carries out an experiment to identify an unknown lead oxide, which is one of the four oxides of lead shown above.

The student plans to reduce the unknown lead oxide to lead by heating the lead oxide in a stream of methane gas, CH_4. The apparatus is shown below.



Student's method

- Weigh an empty dish.
  Add the lead oxide to the dish and reweigh.
- Set up the apparatus and pass methane gas through the tube as shown.
  Heat the dish for 10 minutes.
- Pass cold air through the tube to cool the dish and contents.
- Weigh the dish and contents.

(i) Write the equation for the reduction of Pb_2O_3 with CH_4.

$4Pb_2O_3 + 3CH_4 \rightarrow 8Pb + 3CO_2 + 6H_2O$ **[1]**

(ii) The student uses safety glasses and a lab coat.

State, with a reason, **one** other important safety precaution the student should take when carrying out this experiment.

Compounds may be toxic/poisonous/flammable so use a fume cupboard.
Lead is toxic/poisonous so wear **[1]**
gloves.
Methane is flammable so keep away from flame.

(iii) The student was not sure that all the oxygen had been removed from the lead oxide.

Suggest **two** modifications that the student could make to their method to be confident that all the oxygen had been removed. Explain your reasoning.

1 Heat to a constant mass

pass methane through tube as it cools

2 Spread / break up lead oxide

use excess methane

Bubble gas through lime water. **[2]**

(iv) The student makes suitable modifications to the method and repeats the experiment to obtain the accurate results shown below.

Mass of dish / g	8.364
Mass of dish + lead oxide / g	11.818
Mass of dish + lead at end of experiment / g	11.496

Calculate the empirical formula of the lead oxide.

Pb : O

3.132 : 0.322

$\frac{3.132}{207.2}$: $\frac{0.322}{16}$

0.0151 : 0.020125

3 : 4

11.818 − 11.496 = 0.322

11.496 − 8.364 = 3.132

empirical formula = Pb_3O_4 **[2]**

(b) SiO_2 and CO_2 are oxides of other Group 14 (Group 4) elements.

Solid SiO_2 melts at 2156 °C. Solid CO_2 melts at −56 °C.

Suggest the type of lattice structure in solid SiO_2 and in solid CO_2 and explain the difference in melting points in terms of the types of force within each lattice structure.

Structure in SiO_2(s) giant

Structure in CO_2(s) Simple molecular

Explanation london forces in CO_2. Covalent bonds in SiO_2 stronger than intermolecular forces in CO_2

[4]

10. This question is about magnesium, bromine and magnesium bromide.

(a) Relative atomic mass is defined as 'the weighted mean mass compared with 1/12th mass of carbon-12'.

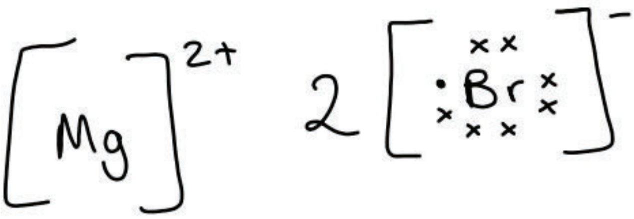
Explain what is meant by the term **weighted mean mass**.

The mean taking into account the relative abundancies of the isotopes

[1]

(b) (i) Draw a 'dot-and-cross' diagram for $MgBr_2$. ionic bonding

Show outer electron shells only.



[2]

(ii) Calculate the total number of **ions** in 1.74 g of magnesium bromide, $MgBr_2$.

Give your answer to **3** significant figures.

$\frac{1.74}{24.3 + (2 \times 79.9)} = 9.45 \times 10^{-3}$ mol

3 ions: Mg^{2+}, Br^-, Br^-

$9.45 \times 10^{-3} \times 3 = 0.0283$ mol

$0.0283 \times 6.023 \times 10^{23} = 1.71 \times 10^{22}$ ions

Avogadro's constant

number of ions = 1.71×10^{22} **[3]**

(c)* **Table 16.1** shows some physical properties of magnesium, bromine and magnesium bromide.

Substance	Melting point / °C	Electrical conductivity	
		Solid	Liquid
Magnesium	711	Good	Good
Bromine	−7	Poor	Poor
Magnesium bromide	650	Poor	Good

Table 16.1

Explain the physical properties shown in **Table 16.1** using your knowledge of structure and bonding. **[6]**

Mg: giant lattice, metallic bonding with delocalised electrons so can conduct

Br: simple molecular, london forces (induced dipole dipole interactions between molecules so doesn't conduct

$MgBr_2$: giant lattice, ionic bonding between oppositely charged ions in solid ions can't move in solution ions can move and conduct

Metallic and ionic bonds are stronger than london forces so the melting points of Mg and $MgBr_2$ are greater.

Additional answer space if required

....

....

....

....

....

....

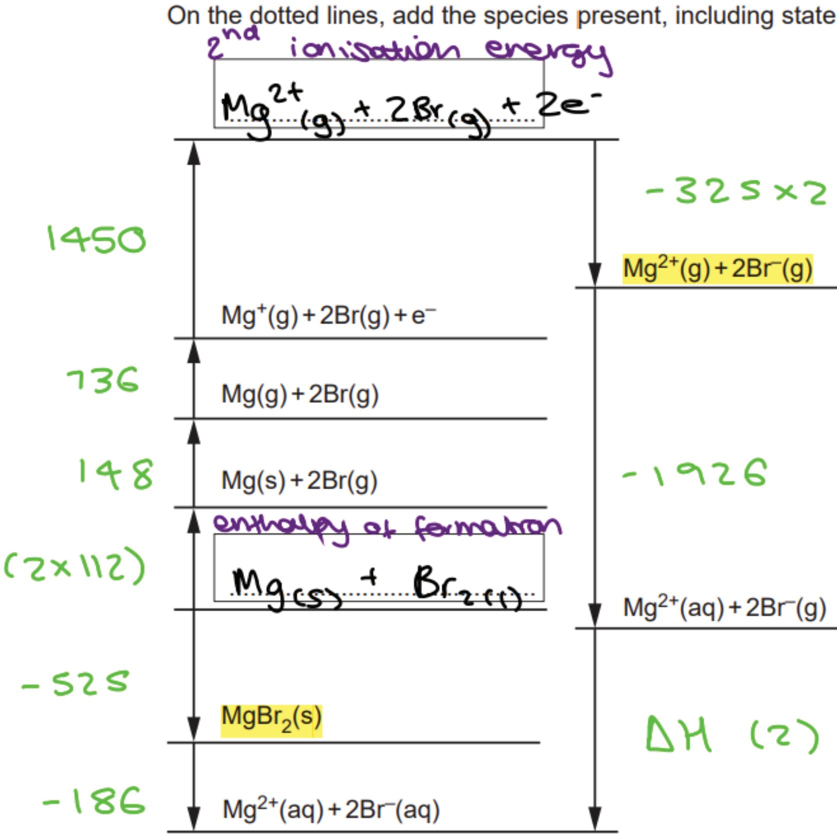
(d) The enthalpy change of hydration of bromide ions can be determined using the enthalpy changes in **Table 16.2**.

Enthalpy change	Energy / kJ mol^{-1}
1st ionisation energy of magnesium	+736
2nd ionisation energy of magnesium	+1450
atomisation of bromine	+112
atomisation of magnesium	+148
electron affinity of bromine	−325
formation of magnesium bromide	−525
hydration of bromide ion	to be calculated
hydration of magnesium ion	−1926
solution of magnesium bromide	−186

Table 16.2

(i) An incomplete energy cycle based on **Table 16.2** is shown below.

On the dotted lines, add the species present, including state symbols.



[2]

(ii) Using your completed energy cycle in **16(d)(i)**, calculate the enthalpy change of hydration of bromide ions.

$1926 + (2 \times 325) - 1450 - 736 - 148 - (2 \times 112) - 525 - 186 = -693$ kJmol^{-1}

$\frac{-693}{2} = -346.5$ kJmol^{-1}

enthalpy change of hydration = -346.5 kJ mol^{-1} **[2]**

(iii) Write the equation for the lattice enthalpy of magnesium bromide and calculate the lattice enthalpy of magnesium bromide.

Equation $Mg^{2+}(g) + 2Br^-(g) \rightarrow MgBr_2(s)$

Calculation

$-1926 + (2 \times -346.5) + 186 = -2433$ kJmol^{-1}

lattice enthalpy = -2433 kJ mol^{-1} **[3]**

11. Which statement about bonds is correct?

A The C=C bond in ethene is ~~more~~ less polar than the C−C bond in ethane.

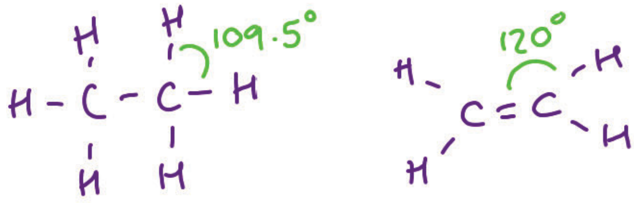
B A σ-bond is stronger than a π-bond. ✓ greater + stronger orbital overlap

C The H−C−H bond angle in ethane is greater than the H−C−H bond angle in ethene.

D A σ-bond is formed from ~~sideways~~ direct overlap of p orbitals.

Your answer B

[1]



120° > 109.5°

12. This question is about two different types of acid found in organic compounds, carboxylic acids and sulfonic acids, as shown in **Fig. 6.1**.

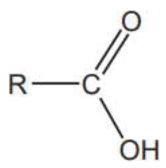
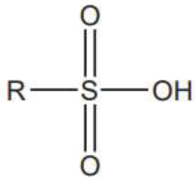
	
Carboxylic acid	Sulfonic acid

Fig. 6.1

(a) Complete **Table 6.1** to predict bond angles ***a*** and ***b*** and name the shapes which makes these bond angles in the functional groups of carboxylic acids and sulfonic acids.

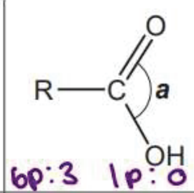
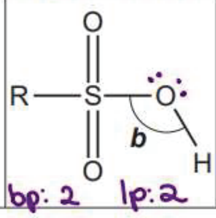
Type of acid	Acid	Bond angle	Name of shape
Carboxylic acid		120°	trigonal planar
Sulfonic acid		109.5°	non-linear

Table 6.1

lone pairs repel more than bonded pairs **[2]**

(b) Ethanoic acid, CH_3COOH, and methanesulfonic acid, CH_3SO_2OH, are both monobasic acids. The pK_a values are shown in the table.

pka = pH

Acid		pK_a
Ethanoic acid	CH_3COOH	4.76
Methanesulfonic acid	CH_3SO_2OH	–1.90

A student suggests that 1.0 mol dm^{-3} CH_3SO_2OH should have a lower pH value than 1.0 mol dm^{-3} CH_3COOH.

Write an equation, showing conjugate acid–base pairs, for the equilibrium of CH_3SO_2OH with water and explain, with reasons, whether the student is correct.

Label the conjugate acid–base pairs: **A1**, **B1** and **A2**, **B2**.

$CH_3SO_2OH + H_2O \rightleftharpoons CH_3SO_2O^- + H_3O^+$

A1 B2 B1 A2

acids: proton donors

bases: proton acceptors

CH_3SO_2OH is a stronger acid / dissociates more

student is correct CH_3SO_2OH has a lower **[4]**

pka / pH / higher ka / [H⁺]

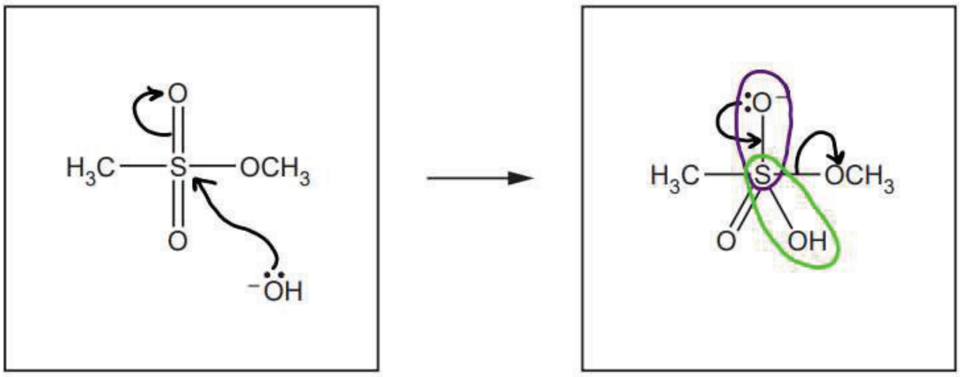
(c) Carboxylic acids and sulfonic acids both form esters.

Sulfonic acid esters can be hydrolysed by aqueous alkali.
The equation shows the alkaline hydrolysis of a sulfonic acid ester.

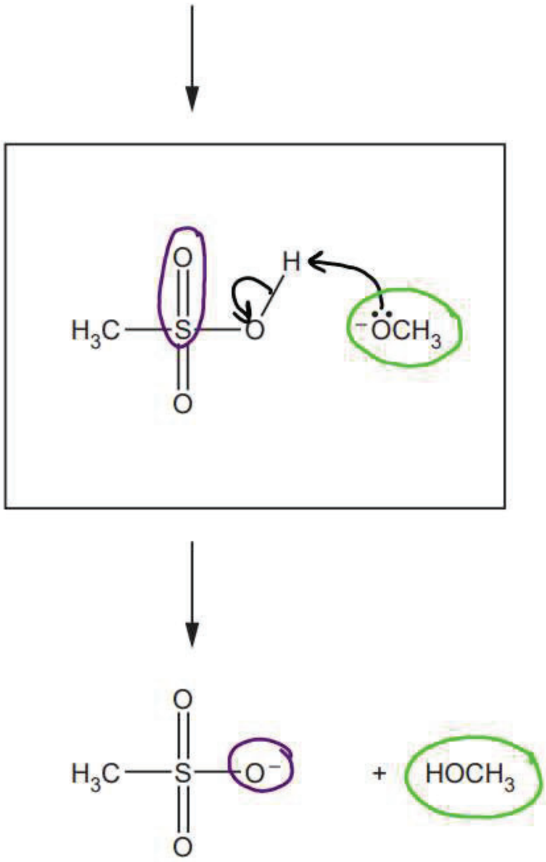
$CH_3SO_2OCH_3 + OH^- \rightarrow CH_3SO_2O^- + CH_3OH$

In the **3 boxes below**, add curly arrows to show the mechanism for this reaction.

In the first box, the hydroxide ion acts as a nucleophile.



look at what you want to make at each stage and work backwards to draw curly arrows



[4]