

# Interactive Example Candidate Responses

## Paper 42 (May/June 2016), Question 2

### Cambridge International AS & A Level Chemistry 9701

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2 Ethanoic acid is a weak acid.

(a) Explain what is meant by the term weak acid.

*It is an acid that does not dissociate completely in a solution.* [1]

(b) The  $pK_a$  values of four acids are listed below.

acid	structural formula	$pK_a$
1	$CH_3CO_2H$	4.8
2	$CH_3CH_2CO_2H$	4.9
3	$CH_3CHClCO_2H$	2.8
4	$CH_2ClCH_2CO_2H$	4.0

(i) State the mathematical relationship between  $pK_a$  and the acid dissociation constant  $K_a$ .

*$pK_a = -\log_{10}(K_a)$*  [1]

(ii) With reference to acidity, explain the difference in  $pK_a$  values between

- acid 1 and acid 2,

*acid 2 is less acidic than acid 1 because of its more powerful inductive effect. Therefore its  $K_a$  is lower and its  $pK_a$  is greater.*

- acid 2 and acid 3,

*acid 2 is less acidic than acid 3 because the chlorine atom is electron withdrawing and weakens the O-H bond increasing  $K_a$  and decreasing  $pK_a$ .*

- acid 3 and acid 4.

*acid 3 is more acidic than acid 4, because the chlorine atom is closer to the carboxylic acid group and therefore has higher  $K_a$  and lower  $pK_a$ .* [3]

Your  
Mark

2(a)

2(b)(i)

2(b)(ii)

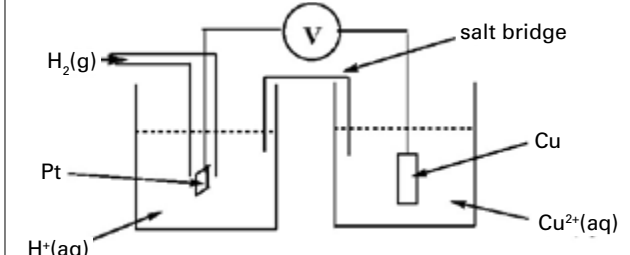
2(c)(i)

2(c)(ii)

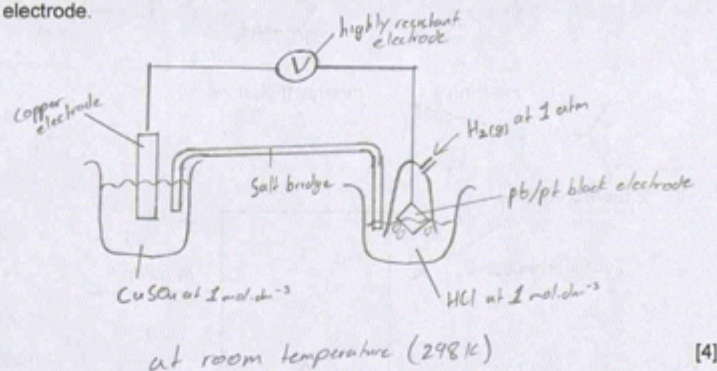
2(d)(i)

2(d)(ii)

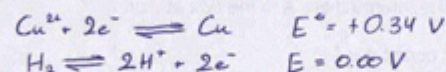
## Q2 Mark scheme

(a)	(an acid that is) partially / incompletely ionised / dissociated [1]
(b)(i)	$pK_a = -\log K_a$ or $K_a = 10^{-pK_a}$ [1]
(b)(ii)	ethanoic acid (1) is <b>more acidic</b> than propanoic acid (2) due to smaller electron-donating (R / alkyl) group / less electron-donating (R / alkyl) group(s) 2-chloropropanoic acid (3) is more acidic than propanoic acid (2) due to electron-withdrawing / electronegative (Cl / chlorine) atom 2-chloropropanoic acid (3) is more acidic than 3-chloropropanoic acid (4) since the Cl / chlorine / electronegative atom is closer to the $CO_2^-$ / acid [3]
(c)(i)	 <p>M1: voltmeter / V <b>and</b> salt bridge labelled M2: Cu and <math>Cu^{2+}</math> / <math>CuSO_4</math> (any soluble Cu(II) salt) M3: <math>H_2</math> (arrow in) <b>and</b> <math>H^+</math> / <math>HCl</math> / <math>H_2SO_4</math> / any mineral acid M4 Pt <b>and</b> one solution at 1 M / 1 mol dm<sup>-3</sup> OR <math>H_2</math> at 1 atm [4]</p>
(c)(ii)	$E^\circ_{cell} = 0.34$ (V) <b>and</b> $(Cu^{2+}) / Cu$ is the positive electrode [3]
(d)(i)	$K_a = 1.23 \times 10^{-5}$ $[H^+] = \sqrt{(K_a \cdot c)} = \sqrt{(1.23 \times 10^{-5} \times 0.1)} = 1.11 \times 10^{-3}$ mol dm <sup>-3</sup> $pH = 3.0$ (2.96) ecf from $[H^+]$ [1]
(d)(ii)	$E = 0.0 + 0.059 \log(1.11 \times 10^{-3})$ OR $= -0.17(4)$ V so new $E_{cell} = 0.34 + 0.17 = 0.51$ V ecf from (d)(i) [2]
[Total: 14]	

- (c) (i) Draw a fully labelled diagram of the equipment needed to measure the voltage of an electrochemical cell consisting of the standard hydrogen electrode and the standard Cu/Cu<sup>2+</sup> electrode.



- (ii) For the cell drawn in (i), calculate the  $E^\circ_{\text{cell}}$  and state which electrode is positive.



$E^\circ_{\text{cell}} = +0.34$  identity of the positive electrode Copper electrode [1]

- (d) A monobasic acid, D, has  $K_a = 1.23 \times 10^{-5} \text{ mol dm}^{-3}$ .

- (i) Calculate the pH of a  $0.100 \text{ mol dm}^{-3}$  solution of D.

$$K_a = \frac{[\text{H}^+]^2}{[\text{D}]}, \quad [\text{H}^+] = \sqrt{[\text{D}] \times K_a} \quad \text{pH} = -\log_{10}([\text{H}^+])$$

$$= \sqrt{(0.1)(1.23 \times 10^{-5})} \quad = -\log_{10}(1.11 \times 10^{-3})$$

$$= 1.11 \times 10^{-3} \quad = 2.95$$

pH = 2.95 [2]

- (ii) An electrochemical cell similar to the one you have drawn in (c)(i) was set up using a  $0.100 \text{ mol dm}^{-3}$  solution of D in the hydrogen electrode instead of the standard solution.

Use the data and the Nernst equation,  $E = E^\circ + 0.059 \log [\text{H}^+(\text{aq})]$ , to calculate the new  $E_{\text{cell}}$  in this experiment.

$$E_{\text{cell}} = E^\circ_{\text{reduced}} - E^\circ_{\text{oxidised}}$$

$$E_{\text{cell}} = 0.34 - (0 + 0.059 \log(1.11 \times 10^{-3}))$$

$$= 0.34 - (-0.17) = 0.51 \text{ V}$$

$E_{\text{cell}} = 0.51 \text{ V}$  [2]

[Total: 14]

Your  
Mark

2(a)

2(b)(i)

2(b)(ii)

2(c)(i)

2(c)(ii)

2(d)(i)

2(d)(ii)

Q2	Mark scheme	
(a)	(an acid that is) partially / incompletely ionised / dissociated	[1]
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(c)(i)	<p>M1: voltmeter / V <b>and</b> salt bridge labelled M2: Cu and Cu<sup>2+</sup> / CuSO<sub>4</sub> (any soluble Cu(II) salt) M3: H<sub>2</sub> (arrow in) <b>and</b> H<sup>+</sup> / HCl / H<sub>2</sub>SO<sub>4</sub> / any mineral acid M4 Pt <b>and</b> one solution at 1 M / 1 mol dm<sup>-3</sup> OR H<sub>2</sub> at 1 atm</p>	[4]
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(d)(ii)	$E = 0.0 + 0.059 \log(1.11 \times 10^{-3})$ OR $= -0.17(4) \text{ V}$ so new $E_{\text{cell}} = 0.34 + 0.17 = \mathbf{0.51 \text{ V}}$ ecf from (d)(i)	[2] [Total: 14]



2 Ethanoic acid is a weak acid.

(a) Explain what is meant by the term weak acid:

Acid which dissociates partially in a solution to form a low concentration of  $H^+$  ions. [1]

(b) The  $pK_a$  values of four acids are listed below.

acid	structural formula	$pK_a$
1	$CH_3CO_2H$	4.8
2	$CH_3CH_2CO_2H$	4.9
→ 3	$CH_3CHClCO_2H$	2.8
4	$CH_2ClCH_2CO_2H$	4.0

(i) State the mathematical relationship between  $pK_a$  and the acid dissociation constant  $K_a$ .

$pK_a = -\lg(K_a)$ .  $-\log_{10} K_a$ . [1]

(ii) With reference to acidity, explain the difference in  $pK_a$  values between

- acid 1 and acid 2,

Ethyl groups of acid 2 is more electron releasing than methyl group of acid 1. Bonding between  $CH_3CH_2CO_2^-$  and  $H^+$  is stronger and so, acid 2 dissociates less extensively than acid 1.

- acid 2 and acid 3.

Acid 3 has an electronegative atom (chlorine) which is electron withdrawing, stabilizing the anion of acid 3. Acid 3 ionizes more extensively than acid 2.

- acid 3 and acid 4.

The position of chlorine atom in acid 3 is at carbon atom adjacent to  $CO_2^-$  group while for acid 4, the chlorine atom is further from  $CO_2^-$  group. Anion of acid 3 is more stable than that of acid 4. [3]

Your  
Mark

2(a)

2(b)(i)

2(b)(ii)

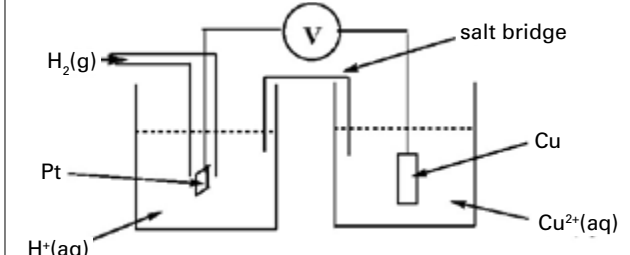
2(c)(i)

2(c)(ii)

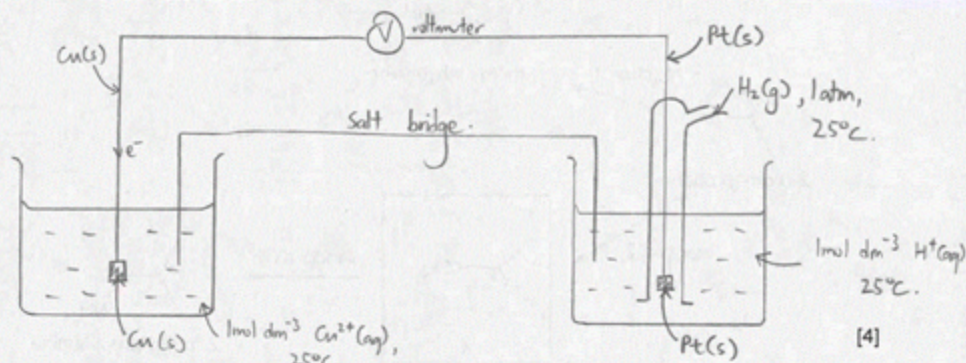
2(d)(i)

2(d)(ii)

## Q2 Mark scheme

(a)	(an acid that is) partially / incompletely ionised / dissociated [1]
(b)(i)	$pK_a = -\log K_a$ or $K_a = 10^{-pK_a}$ [1]
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(c)(i)	 <p>M1: voltmeter / V <b>and</b> salt bridge labelled M2: Cu and <math>Cu^{2+}</math> / <math>CuSO_4</math> (any soluble Cu(II) salt) M3: <math>H_2</math> (arrow in) <b>and</b> <math>H^+</math> / <math>HCl</math> / <math>H_2SO_4</math> / any mineral acid M4 Pt <b>and</b> one solution at 1 M / 1 mol dm<sup>-3</sup> OR <math>H_2</math> at 1 atm [4]</p>
(c)(ii)	$E^\circ_{cell} = 0.34$ (V) <b>and</b> ( $Cu^{2+}$ ) / Cu is the positive electrode [3]
(d)(i)	$K_a = 1.23 \times 10^{-5}$ $[H^+] = \sqrt{K_a \cdot c} = \sqrt{(1.23 \times 10^{-5} \times 0.1)} = 1.11 \times 10^{-3}$ mol dm <sup>-3</sup> $pH = 3.0$ (2.96) ecf from $[H^+]$ [1]
(d)(ii)	$E = 0.0 + 0.059 \log(1.11 \times 10^{-3})$ OR $= -0.17(4)V$ so new $E_{cell} = 0.34 + 0.17 = 0.51V$ ecf from (d)(i) [2] [Total: 14]

- (c) (i) Draw a fully labelled diagram of the equipment needed to measure the voltage of an electrochemical cell consisting of the standard hydrogen electrode and the standard Cu/Cu<sup>2+</sup> electrode.



- (ii) For the cell drawn in (i), calculate the  $E_{\text{cell}}^{\circ}$  and state which electrode is positive.

$E_{\text{cell}}^{\circ} = +0.34 \text{ V}$  identity of the positive electrode Platinum (hydrogen electrode) [1]

- (d) A monobasic acid, D, has  $K_{\text{a}} = 1.23 \times 10^{-5} \text{ mol dm}^{-3}$ .

- (i) Calculate the pH of a  $0.100 \text{ mol dm}^{-3}$  solution of D.

$$[\text{H}^{+}] = \sqrt{(0.100 \times 1.23 \times 10^{-5})} = 1.09 \times 10^{-3} \text{ mol dm}^{-3}$$

$$\therefore \text{pH} = -\lg(1.09 \times 10^{-3}) = 2.96$$

$\text{pH} = 2.96$  [2]

- (ii) An electrochemical cell similar to the one you have drawn in (c)(i) was set up using a  $0.100 \text{ mol dm}^{-3}$  solution of D in the hydrogen electrode instead of the standard solution.

Use the data and the Nernst equation,  $E = E^{\circ} + 0.059 \log [\text{H}^{+}(\text{aq})]$ , to calculate the new  $E_{\text{cell}}$  in this experiment.

$$E = 0.34 + 0.059 \lg(1.09 \times 10^{-3})$$

$$= 0.166 \text{ V}$$

$E_{\text{cell}} = 0.166 \text{ V}$  [2]

[Total: 14]

Your  
Mark

2(a)

2(b)(i)

2(b)(ii)

2(c)(i)

2(c)(ii)

2(d)(i)

2(d)(ii)

Q2	Mark scheme	
(a)	(an acid that is) partially / incompletely ionised / dissociated	[1]
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(c)(ii)	$E^{\circ}_{\text{cell}} = 0.34 \text{ (V)}$ <b>and</b> (Cu <sup>2+</sup> ) / Cu is the positive electrode	[3]
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(d)(ii)	$E = 0.0 + 0.059 \log(1.11 \times 10^{-3})$ OR $= -0.17(4) \text{ V}$ so new $E_{\text{cell}} = 0.34 + 0.17 = 0.51 \text{ V}$ ecf from (d)(i)	[2] [Total: 14]

2 Ethanoic acid is a weak acid.

(a) Explain what is meant by the term weak acid.

~~Weak acid produces the products less  $H^+$  ions~~  
Weak acid dissociates to produce ~~less~~  $H^+$  ions and doesn't ionise completely. [1]

(b) The  $pK_a$  values of four acids are listed below.

acid	structural formula	$pK_a$
1	$CH_3CO_2H$	4.8
2	$CH_3CH_2CO_2H$	4.9
3	$CH_3CHClCO_2H$	2.8
4	$CH_2ClCH_2CO_2H$	4.0

(i) State the mathematical relationship between  $pK_a$  and the acid dissociation constant  $K_a$ .

$pK_a = -\log_{10} K_a$  [1]

(ii) With reference to acidity, explain the difference in  $pK_a$  values between

- acid 1 and acid 2,

Acid 2 is a <sup>slightly</sup> weak acid compared to acid 1  
as the Acid 2  $pK_a$  value is higher than that of Acid 1  
as ethyl group is electron withdrawing group

- acid 2 and acid 3,

higher the  $pK_a$  value, lesser the acidity. Therefore, Acid 3  
is a stronger acid than Acid 2 due to the ethyl group present  
in Acid 2 and Cl in Acid 3 which is electron donating group.

- acid 3 and acid 4.

Acid 3 is ~~the~~ acidic than Acid 4 as there is  
present in Acid 3. [3]

Your  
Mark

2(a)

2(b)(i)

2(b)(ii)

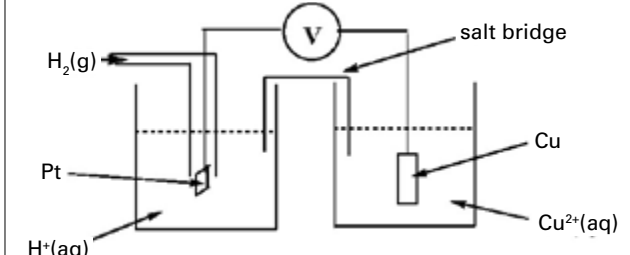
2(c)(i)

2(c)(ii)

2(d)(i)

2(d)(ii)

## Q2 Mark scheme

(a)	(an acid that is) partially / incompletely ionised / dissociated [1]
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(d)(i)	$K_a = 1.23 \times 10^{-5}$ $[H^+] = \sqrt{(K_a \cdot c)} = \sqrt{(1.23 \times 10^{-5} \times 0.1)} = 1.11 \times 10^{-3}$ mol dm <sup>-3</sup> $pH = 3.0$ (2.96) ecf from $[H^+]$ [1]
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Your  
Mark

2(a)

2(b)(i)

2(b)(ii)

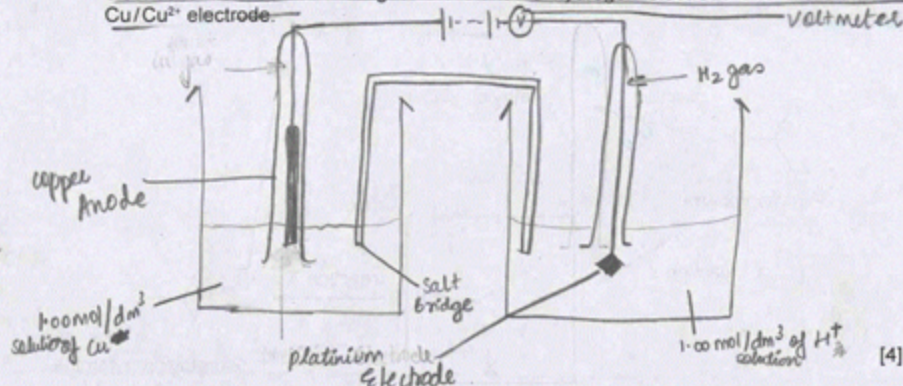
2(c)(i)

2(c)(ii)

2(d)(i)

2(d)(ii)

- (c) (i) Draw a fully labelled diagram of the equipment needed to measure the voltage of an electrochemical cell consisting of the standard hydrogen electrode and the standard Cu/Cu<sup>2+</sup> electrode.



- (ii) For the cell drawn in (i), calculate the  $E^\circ_{\text{cell}}$  and state which electrode is positive.

$E^\circ_{\text{cell}} = \dots \dots \dots +0.34$  identity of the positive electrode  $\dots \dots \dots$  copper [1]

- (d) A monobasic acid, D, has  $K_a = 1.23 \times 10^{-5} \text{ mol dm}^{-3}$ .

- (i) Calculate the pH of a  $0.100 \text{ mol dm}^{-3}$  solution of D.

$pH = pK_a + \log_{10} \left[ \frac{\text{salt}}{\text{acid}} \right]$   
 $pK_a = -\log_{10} 1.23 \times 10^{-5} = -8.99 \times 10^{-7}$   
 $pH = -\log_{10} [H^+] = -\log_{10} 1.23 \times 10^{-5}$   
 $pH = \dots \dots \dots$  [2]

- (ii) An electrochemical cell similar to the one you have drawn in (c)(i) was set up using a  $0.100 \text{ mol dm}^{-3}$  solution of D in the hydrogen electrode instead of the standard solution.

Use the data and the Nernst equation,  $E = E^\circ + 0.059 \log [H^+(aq)]$ , to calculate the new  $E_{\text{cell}}$  in this experiment.

$E = 0.34 + 0.059 \log \sqrt{0.100}$   
 $= 0.3105$   
 $E_{\text{cell}} = \dots \dots \dots 0.3105 \text{ V}$  [2]

[Total: 14]

Q2	Mark scheme	
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(d)(ii)	$E = 0.0 + 0.059 \log (1.11 \times 10^{-3})$ OR $= -0.17(4) \text{ V}$ so new $E_{\text{cell}} = 0.34 + 0.17 = \mathbf{0.51 \text{ V}}$ ecf from (d)(i)	[2] [Total: 14]



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