2

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91166

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QUALIFY FOR THE FUTURE WORLD
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Level 2 Chemistry, 2017

91166 Demonstrate understanding of chemical reactivity

2.00 p.m. Thursday 16 November 2017
Credits: Four

Achievement	Achievement with Merit	Achievement with Excellence
Demonstrate understanding of chemical reactivity.	Demonstrate in-depth understanding of chemical reactivity.	Demonstrate comprehensive understanding of chemical reactivity.

Check that the National Student Number (NSN) on your admission slip is the same as the number at the top of this page.

You should attempt ALL the questions in this booklet.

A periodic table is provided on the Resource Sheet L2–CHEMR.

If you need more room for any answer, use the extra space provided at the back of this booklet and clearly number the question.

Check that this booklet has pages 2–12 in the correct order and that none of these pages is blank.

YOU MUST HAND THIS BOOKLET TO THE SUPERVISOR AT THE END OF THE EXAMINATION.

Excellence

TOTAL 23

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4 A4 1×10^-4

QUESTION ONE

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(a) Propanoic acid, C_2H_5COOH, is dissolved in water and the resulting solution has a pH of 4.2.

(i) Complete the equation by writing the formulae of the two products.

$$C_2H_5COOH(aq) + H_2O(\ell) \rightleftharpoons C_2H_5COO^- + H_3O^+$$

(ii) Explain the proton, H^+, transfer in this reaction, and identify the two conjugate acid-base pairs.

C_2H_5COOH donates a proton to H_2O to form $C_2H_5COO^-$ whilst H_2O forms H_3O^+. This means that C_2H_5COOH is a conjugate acid as it donates a proton to H_2O. This forms its conjugate base $C_2H_5COO^-$. H_2O acts as a base in this instance accepting a proton to form its conjugate acid H_3O^+.

(b) Sodium ethanoate, $CH_3COONa(s)$, is a salt. When dissolved in water, it dissociates into ions.

Explain, including TWO relevant equations, whether a solution of sodium ethanoate is acidic or basic.

$CH_3COONa \xrightarrow{H_2O} CH_3COO^- + Na$

$CH_3COO^- + H_2O \rightleftharpoons CH_3COOH + OH^-$

It is a basic solution as it first produces CH_3COO^- ions which then dissociate partially in water to form OH^- ions meaning the solution must be basic.

(c) (i) A solution of sodium hydroxide, NaOH(*aq*), has a pH of 11.6.

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Calculate the hydronium ion concentration $[H_3O^+]$, and the hydroxide ion concentration, $[OH^-]$, in the solution.

$K_w = 1 \times 10^{-14}$

$[H_3O^+]$ = 10^{-pH}

$[H_3O^+] = 10^{-11.6}$

$[H_3O^+] = 2.51 \times 10^{-12}$ (3 s.f.) mol L^{-1} (3.s.f)

$[OH^-]$ = $[OH^-] = \frac{1 \times 10^{-14}}{2.51 \times 10^{-12}}$

$[OH^-]$ = 0.00398 molL^{-1} (3.s.f)

(ii) Calculate the pH of a 2.96×10^{-4} mol L^{-1} solution of potassium hydroxide, KOH(*aq*).

pH = $-\log(2.96 \times 10^{-4})$

pH = 3.528   pH = 14 − 3.528   pH = 10.5 (3.s.f)

(d) Solutions of ammonia, $NH_3(aq)$, and sodium carbonate, $Na_2CO_3(aq)$, are both basic.

Compare and contrast the electrical conductivity of these two solutions.

Na_2CO_3 is a good electrical conductor this is because in solution it dissociates into ions with the following equation: $Na_2CO_3 + H_2O \rightarrow 2Na^+ + CO_3^{2-}$ This means there are alot of charged ions in solution that are free to move. Therefore it is a good electrical conductor.

NH_3 however is not a good electrical conductor. It does not fully dissociate in water and has the following equation: $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ Therefore, there is a relatively low concentration of OH^- ions and NH_4^+ ions. It therefore conducts very poorly compared with Na_2CO_3 as it has a far lower concentration of free moving ions in solution.

E8

QUESTION TWO

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The addition of a small amount of iron to a mixture of nitrogen and hydrogen gases helps to speed up the production of ammonia gas.

$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

(a) Identify and explain the role of iron in this reaction.

In your answer, you should refer to activation energy and collision theory.

You may include a diagram or diagrams in your answer.

Iron is the catalyst in this reaction. A catalyst ~~acts~~ forms an alternate reaction pathway lowering the amount of activation energy required for the reaction to take place. This means that when reactant particles collide more collisions will have enough energy to overcome this activation energy barrier resulting in a successful collision. Thus, there will be more succesful collisions per second and the rate of reaction will be far greater than a reaction without a catalyst. ~~This will mean~~ However, this will not mean more product is formed only that the rate of reaction will be faster than without a catalyst.

The reaction described above is an equilibrium reaction, as represented by the following equation:

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

(b) (i) Write the equilibrium constant expression for this reaction.

$K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$

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(ii) The value of the equilibrium constant, K_c, is 640 at 25°C.

Show, by calculation, using the concentrations of the gases given in the table below, whether or not the reaction is at equilibrium.

Explain your answer.

Gas	N_2	H_2	NH_3
Concentration (mol L⁻¹)	0.0821	0.0583	0.105

Is the mixture at equilibrium?

(Circle) **Yes** **No**

Calculation and explanation: let kc = Q

$Q = \frac{[NH_3]^2}{[N_2][H_2]^3}$ $Q = \frac{[0.105]^2}{[0.0821] \times [0.0583]^3}$

Q = 677.608

Q = 678 (3 s.f.) Q ≠ kc

Q is not equal to kc therefore the equilibrium is not at equilibrium and therefore one side of the reaction will be favoured.

(c) As the temperature increases, the value of the equilibrium constant, K_c, decreases from 640 at 25°C to 0.440 at 200°C.

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Justify whether the formation of ammonia, $NH_3(g)$, is an endothermic or exothermic reaction.

When the temperature is increased the equilibrium will shift to minimise the change that is being imposed on it. Therefore it will favour the endothermic reaction to use up this added heat energy. When the temperature was increased kc decreased. To reduce kc the reverse reaction must be favoured as $kc = \frac{[products]}{[reactants]}$ therefore the reactants must be at a higher concentration to reduce kc and the reverse reaction is favoured. Therefore, if kc decreased when the temperature increased the reverse reaction must be an endothermic one. If the reverse reaction is endothermic the forward reaction must be exothermic. Thus, the formation of NH_3 is exothermic as it is the forward reaction.

E7

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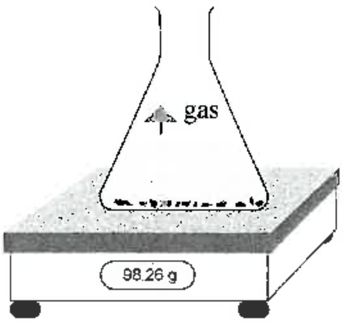
QUESTION THREE

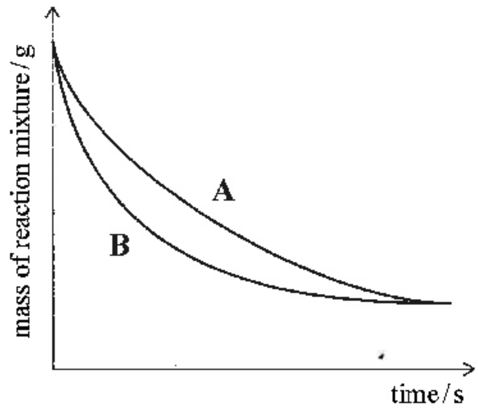
(a) Consider the reaction between calcium carbonate powder, $CaCO_3(s)$, and a solution of hydrochloric acid, $HCl(aq)$.

As the reaction proceeds, the mass of the reaction mixture decreases as carbon dioxide gas, $CO_2(g)$, escapes.

This is represented on the graph below.

Line A represents the reaction occurring at 20°C and line B represents the reaction occurring at 40°C.





Compare and contrast the reaction between calcium carbonate powder, $CaCO_3(s)$, and a solution of hydrochloric acid, $HCl(aq)$ at two temperatures: 20°C and 40°C, assuming all other conditions are kept the same.

Your answer should refer to collision theory and rates of reaction.

When the temperature is increased the rate of reaction will increase compared to a reaction occuring at a lower temperature. An increase in temperature will mean that reactant particles will have more kinetic energy. Therefore, when particles collide more of the particles will have sufficient energy to overcome the activation energy barrier and result in a effective collisions. There will therefore be more effective collisions per second and thus the rate of reaction will be higher for the mixture at the higher temperature. Temperature also increases the speed of the reactant particles. This will mean a [illegible] collision is far

There is more space for your answer to this question on the following page.

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more likely to occur and there will be more collisions per second. Therefore, there will be more effective collisions per second and the rate of reaction will be greater for a mixture at higher temperature. With a greater rate of reaction the mixture will produce the same amount of CO_2 gas just at a faster rate of reaction if it is heated compared with that of a mixture at a lower temperature.

(b) Two different cobalt(II) complex ions, $[Co(H_2O)_6]^{2+}$ and $[CoCl_4]^{2-}$, exist together in a solution in equilibrium with chloride ions, $Cl^-(aq)$.

The forward reaction is endothermic; ΔH is positive. The equation for this equilibrium is shown below.

$$[Co(H_2O)_6]^{2+}(aq) + 4Cl^-(aq) \rightleftharpoons [CoCl_4]^{2-}(aq) + 6H_2O(\ell)$$

pink blue

Explain using equilibrium principles, the effect on the colour of the solution if:

(i) more water is added to the reaction mixture

If more water is added to the solution will turn pinker. This is because the equilibrium will shift to minimise a change imposed on it. Therefore if more product is added the equilibrium will shift to use it up and shift in the reverse direction making more reactants which are pink.

(ii) a test tube containing the reaction mixture is placed in a beaker of ice-cold water.

If the mixture is cooled the equm solution will become pinker. This is because the equilibrium mixture will shift to minimise the change imposed on it. If it is cooled it will ~~shift to the exoth~~ favour the exothermic reaction to heat up the mixture. As the forward reaction is endothermic the reverse reaction must be exothermic. Therefore it will favour the reverse reaction creating more reactants which are pink and turning the solution a pinker colour.

(c) Brown nitrogen dioxide gas, $NO_2(g)$, exists in equilibrium with the colourless gas, dinitrogen tetroxide, $N_2O_4(g)$.

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$$2NO_2(g) \rightleftharpoons N_2O_4(g)$$

brown colourless

Explain using equilibrium principles, the effect of decreasing the volume of the container (therefore increasing the pressure) on the observations of this equilibrium mixture.

If the pressure of this reaction is increased the equilibrium will try to ~~reduce~~ minimise this change by reducing the pressure. On the left hand side of this reaction there are two gaseous moles and on the right there is only one. The reaction will therefore shift to favour the forward reaction and reduce the pressure on the equilibrium mixture. This means there will be a greater concentration of products to reactants and ~~t~~ as the products are colourless ~~the~~ the ~~so~~ gas will become a ~~cle~~ clearer/lighter colour.

E8

Extra paper if required.
Write the question number(s) if applicable.

QUESTION NUMBER

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Extra paper if required.
Write the question number(s) if applicable.

QUESTION NUMBER

ASSESSOR'S USE ONLY

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QUESTION NUMBER

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Excellence exemplar for 91166 2017			Total score	23
Q	**Grade score**	**Annotation**		
1	E8	The candidate has: completed the equation for propanoic acid reacting with water; identified the correct conjugate acid-base pairs, including an explanation of proton transfer; explained why sodium ethanoate is basic with the relevant equation; calculated $[H_3O^+]$, $[OH^-]$, and the pH of a strong base; and compared and contrasted the electrical conductivity of solutions of ammonia and sodium carbonate.		
2	E7	The candidate has: explained that a catalyst provides an alternative pathway with a lower activation energy, so more collisions have sufficient energy to overcome the activation energy; written the correct equilibrium constant expression; calculated the correct K_c and explained why the reaction is not at equilibrium; and justified why the formation of ammonia is exothermic. If the candidate had explained that the equilibrium shifts in the endothermic direction when the temperature is increased to 'absorb' heat energy rather than 'use up' heat energy, this would have provided evidence towards E8.		
3	E8	The candidate has: explained that at a higher temperature more particles will have sufficient kinetic energy to overcome the activation energy and therefore there will be more effective collisions per second, although the same total mass of carbon dioxide is produced at each temperature; and fully explained the effect of changing the concentration of a product, the temperature, and the pressure on the position of an equilibrium, including expected colour changes.		