



Exam Papers Practice

Boost your performance and confidence with these topic-based exam questions

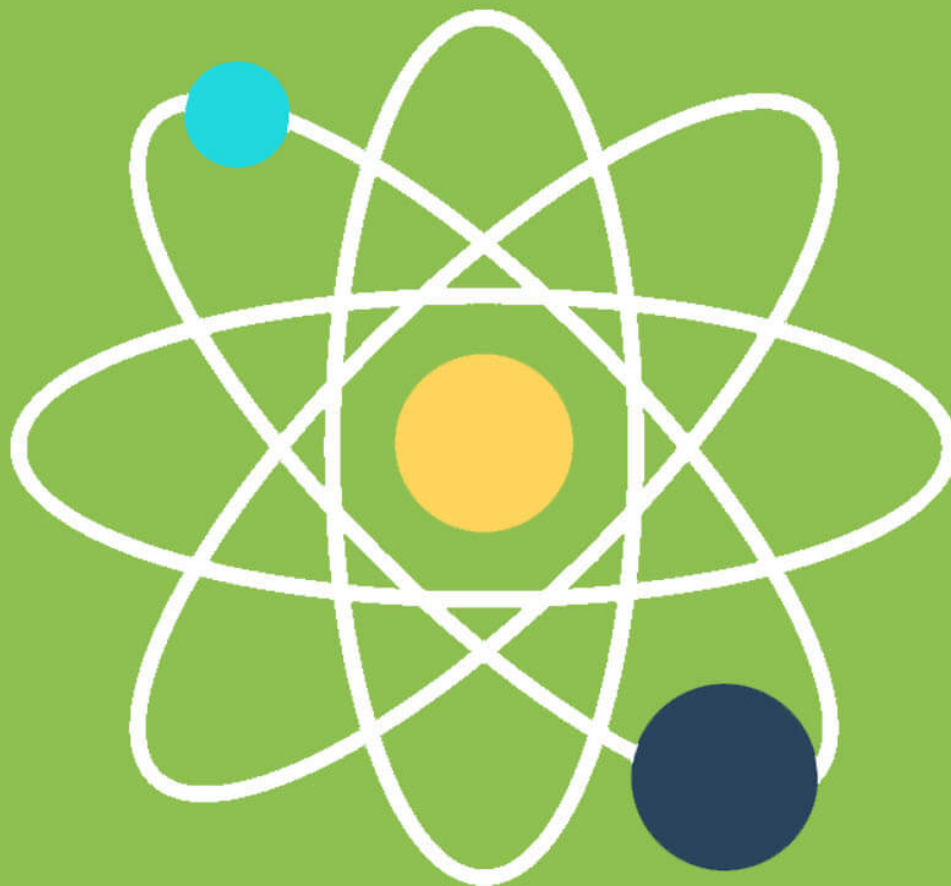
Practice questions created by actual examiners and assessment experts

Detailed mark schemes

Suitable for all boards

Designed to test your ability and thoroughly prepare you

18.2 Calculations Involving Acids & Bases



IB Chemistry - Revision Notes

www.exampaperspractice.co.uk

18.2.1 Acid & Base Dissociation Constants

Acid & Base Dissociation Constants

Weak acids

- A **weak acid** is an acid that **partially** (or incompletely) **dissociates** in aqueous solutions
 - For example, most carboxylic acids (e.g. ethanoic acid), HCN (hydrocyanic acid), H₂S (hydrogen sulfide) and H₂CO₃ (carbonic acid)
 - In general, the following equilibrium is established:



OR



- At equilibrium, the majority of HA molecules remain unreacted
- The position of the equilibrium is more over to the **left** and an equilibrium is established
- As this is an equilibrium, we can write an equilibrium constant expression for the reaction
- This constant is called the **acid dissociation constant**, K_a , and has the units mol dm⁻³

$$K_a = \frac{[\text{A}^-][\text{H}_3\text{O}^+]}{[\text{HA}]}$$

$$K_a = \frac{[\text{A}^-][\text{H}^+]}{[\text{HA}]}$$

Copyright

© 2024 Exam Papers Practice

Acid dissociation constant expressions

- Carboxylic acids are weak acids
 - For example, propanoic acid, CH₃CH₂COOH(aq), dissociates according to the following equation which leads to the K_a expression for propanoic acid:



OR





$$K_a = \frac{[\text{CH}_3\text{CH}_2\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{CH}_2\text{COOH}]}$$

$$K_a = \frac{[\text{CH}_3\text{CH}_2\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{CH}_2\text{COOH}]}$$

Acid dissociation constant expressions for propanoic acid

- Values of K_a are very small
 - For example, K_a for propanoic acid = $1.34 \times 10^{-5} \text{ mol dm}^{-3}$
- When writing the equilibrium expression for weak acids, we assume that the concentration of H_3O^+ (aq) due to the ionisation of water is negligible

Weak bases

- A weak base will also ionise in water and we can represent this with the **base dissociation constant, K_b**
- In general the equilibrium established is:



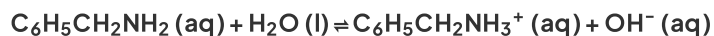
$$K_b = \frac{[\text{BH}^+][\text{OH}^-]}{[\text{B}]}$$

Base dissociation constant expression

Copyright

© 2024 Amines are weak bases

- For example, 1-phenylmethanamine, $\text{C}_6\text{H}_5\text{CH}_2\text{NH}_2$ (aq), dissociates according to the following equation which leads to the K_a expression for 1-phenylmethanamine:



$$K_b = \frac{[\text{C}_6\text{H}_5\text{CH}_2\text{NH}_3^+][\text{OH}^-]}{[\text{C}_6\text{H}_5\text{CH}_2\text{NH}_2]}$$

Base dissociation constant expression for 1-phenylmethanamine

$\text{p}K_a$ and $\text{p}K_b$



- The range of values of K_a and K_b is very wide and for weak acids, the values themselves are very small numbers

Table of K_a values

Acid	$K_a / \text{mol dm}^{-3}$
Methanoic, HCOOH	1.77×10^{-4}
Ethanoic, CH ₃ COOH	1.74×10^{-5}
Benzoic, C ₆ H ₅ COOH	6.46×10^{-5}
Carbonic, H ₂ CO ₃	4.30×10^{-7}

- For this reason, it is easier to work with another term called pK_a for acids or pK_b for bases
- In order to convert the values we need to apply the following calculations:

$$pK_a = -\log K_a \quad K_a = 10^{-pK_a}$$

$$pK_b = -\log K_b \quad K_b = 10^{-pK_b}$$

Table of pK_a values

Acid	$K_a / \text{mol dm}^{-3}$	pK_a
Methanoic, HCOOH	1.77×10^{-4}	3.75
Ethanoic, CH ₃ COOH	1.74×10^{-5}	4.75
Benzoic, C ₆ H ₅ COOH	6.46×10^{-5}	4.18
Carbonic, H ₂ CO ₃	4.30×10^{-7}	6.36

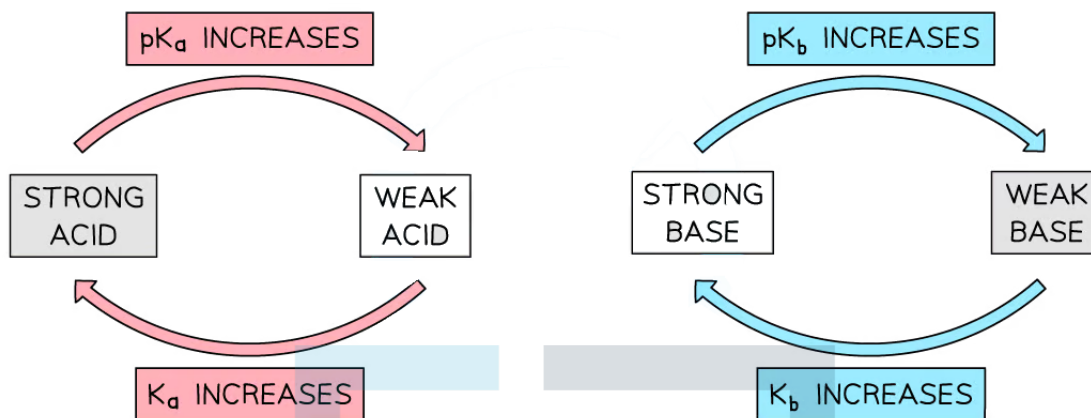
- The range of pK_a values for most weak acids lies between 3 and 7

Relative Strengths of Acids and Bases

- The larger the K_a value, the stronger the acid
- The larger the pK_a value, the weaker the acid



- The larger the K_b value, the stronger the base
- The larger the pK_b value, the weaker the base



pK_a and pK_b tell us the relative strengths of acids and bases

Exam Papers Practice

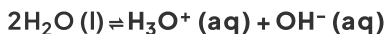
Copyright

© 2024 Exam Papers Practice

Relating Kw to Ka

The Ionic Product of Water and Temperature

- In all aqueous solutions, an equilibrium exists in water where a few water molecules dissociate into protons and hydroxide ions
- We can derive an equilibrium constant for the reaction:



- The concentration of water is constant, so the expression for K_w is:

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

- This is a specific equilibrium constant called the **ionic product for water**
- The product of the two ion concentrations is $1 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$ at 25°C
- For conjugate acid-base pairs, K_a and K_b are related to K_w

$$K_a K_b = K_w$$

- The conjugate base of ethanoic acid is the ethanoate ion, $\text{CH}_3\text{COO}^-(\text{aq})$



acid

conjugate base

- We can then put this in to the K_a expression

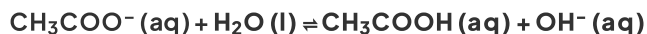
$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]}$$

Copyright

© 2024 Exam Papers Practice

Acid dissociation constant for ethanoic acid

- The ethanoate ion will react with water according to the following equation



- We can then put this in to the K_b expression

$$K_b = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]}$$

**Base dissociation constant for the ethanoate ion**

- Now, these two expressions can be combined, which corresponds to
 - $K_a K_b = K_w$
 - $K_a K_b = 10^{-14}$
- Or we could say that
 - $pK_a + pK_b = pK_w$
 - $pK_a + pK_b = 14$
 - This makes the numbers much more easy to deal with as using $K_a K_b = 10^{-14}$ will give very small numbers

$$K_a K_b = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]} \times \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]}$$

$$= [\text{H}_3\text{O}^+][\text{OH}^-] = K_w$$

Combining $K_a K_b$ expressions

- Or rearranging these:
 - $K_a = K_w / K_b$
 - $K_b = K_w / K_a$

The ionic product of water, K_w

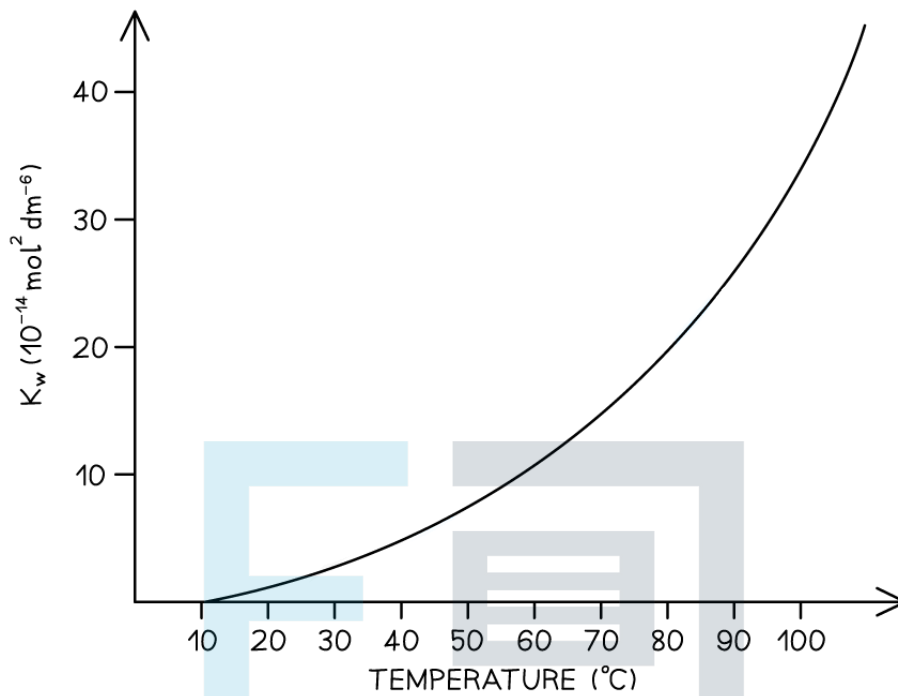
- The ionisation of water is an **endothermic** process



Copyright

© 2024 Exam Papers Practice

- In accordance with Le Châtelier's principle, an increase in temperature will result in the forward reaction being favoured
 - This causes an increase in the concentration of the hydrogen and hydroxide ions
 - This leads to the **magnitude of K_w increasing**
 - Therefore, the **pH will decrease**
- Increasing the temperature, decreases the pH of water (becomes more acidic)
- Decreasing the temperature, increases the pH of water (becomes more basic)



Relationship between K_w and temperature

Exam Papers Practice

Copyright

© 2024 Exam Papers Practice

18.2.2 Acid & Base Problem Solving

Acid & Base Calculations

pH

- The acidity of an aqueous solution depends on the number of H_3O^+ ions in solution
- pH is defined as:

$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

- Where $[\text{H}_3\text{O}^+]$ is the concentration of H_3O^+ ions in mol dm^{-3}
- Similarly, the **concentration of H^+** of a solution can be calculated if the pH is known by rearranging the above equation to:

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

- The pH scale is a logarithmic scale with base 10
- This means that each value is 10 times the value below it
 - For example, pH 5 is 10 times more acidic than pH 6
- pH values are usually given to 2 decimal places

pOH

- The basicity of an aqueous solution depends on the number of hydroxide ions, OH^- , in solution
- pOH is defined as:

$$\text{pOH} = -\log [\text{OH}^-]$$

Copyright

© 2024 Exam Papers Practice

- Where $[\text{OH}^-]$ is the concentration of hydroxide ions in mol dm^{-3}
- Similarly, the **concentration of OH^-** of a solution can be calculated if the pH is known by rearranging the above equation to:

$$[\text{OH}^-] = 10^{-\text{pOH}}$$

- If you are given the concentration of a basic solution and need to find the pH, this can be done by:

$$[\text{H}_3\text{O}^+] = K_w / [\text{OH}^-]$$

- Alternatively, if you are given the $[\text{OH}^-]$ and calculate the pOH, the pH can be found by:

$$\text{pH} = 14 - \text{pOH}$$



Worked example

pH and H_3O^+ calculations

1. Find the pH when the hydrogen ion concentration is $1.60 \times 10^{-4} \text{ mol dm}^{-3}$
2. Find the hydrogen ion concentration when the pH is 3.10

Answers

Answer 1:

The pH of the solution is:

- $\text{pH} = -\log [\text{H}_3\text{O}^+]$
 - $\text{pH} = -\log 1.6 \times 10^{-4}$
 - $\text{pH} = \mathbf{3.80}$

Answer 2:

The hydrogen concentration can be calculated by rearranging the equation for pH

- $\text{pH} = -\log [\text{H}_3\text{O}^+]$
- $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$
 - $[\text{H}_3\text{O}^+] = 10^{-3.10}$
 - $[\text{H}_3\text{O}^+] = \mathbf{7.94 \times 10^{-4} \text{ mol dm}^{-3}}$

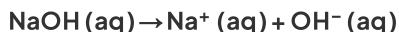
Worked example

pH calculations of a strong alkali

1. Calculate the pH of 0.15 mol dm^{-3} sodium hydroxide, NaOH
2. Calculate the hydroxide concentration of a solution of sodium hydroxide when the pH is 10.50

Answers

Sodium hydroxide is a strong base which ionises as follows:



Answer 1:

The pH of the solution is:



- $[H^+] = K_w \div [OH^-]$
 - $[H^+] = (1 \times 10^{-14}) \div 0.15 = 6.66 \times 10^{-14}$
- $pH = -\log[H^+]$
 - $pH = -\log 6.66 \times 10^{-14} = 13.17$

Answer 2

Step 1: Calculate hydrogen concentration by rearranging the equation for pH

- $pH = -\log[H^+]$
- $[H^+] = 10^{-pH}$
 - $[H^+] = 10^{-10.50}$
 - $[H^+] = 3.16 \times 10^{-11} \text{ mol dm}^{-3}$

Step 2: Rearrange the **ionic product of water** to find the concentration of hydroxide ions

- $K_w = [H^+][OH^-]$
- $[OH^-] = K_w \div [H^+]$

Step 3: Substitute the values into the expression to find the concentration of hydroxide ions

- Since K_w is $1 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$
 - $[OH^-] = (1 \times 10^{-14}) \div (3.16 \times 10^{-11})$
 - $[OH^-] = 3.16 \times 10^{-4} \text{ mol dm}^{-3}$

K_a , pK_a , K_b and pK_b

- In reactions of weak acids and bases, we cannot make the same assumptions as for the ionisation of strong acids and bases
- For a weak acid and its conjugate base, we can use the equation:

$$K_w = K_a K_b$$

- By finding the $-\log$ of these, we can use:

$$pK_w = pK_a + pK_b$$

- Remember, to convert these terms you need to use:

$$pK_a = -\log K_a \quad K_a = 10^{-pK_a}$$

$$pK_b = -\log K_b \quad K_b = 10^{-pK_b}$$

- The assumptions we must make when calculating values for K_a , pK_a , K_b and pK_b are:
 - The initial concentration of acid \approx the equilibrium concentration of acid



- $[A^-] = [H_3O^+]$
- There is negligible ionisation of the water, so $[H_3O^+]$ is not affected
- The temperature is 25 °C

Worked example

Calculate the acid dissociation constant, K_a , at 298 K for a 0.20 mol dm^{-3} solution of propanoic acid with a pH of 4.88.

Answer

Step 1: Calculate $[H_3O^+]$ using

- $[H_3O^+] = 10^{-\text{pH}}$
 - $[H_3O^+] = 10^{-4.88}$
 - $[H_3O^+] = 1.3182 \times 10^{-5}$

Step 2: Substitute values into K_a expression

- $K_a = [H_3O^+]^2 / [CH_3CH_2COOH]$
 - $K_a = (1.3182 \times 10^{-5})^2 / 0.2$
 - $K_a = 8.70 \times 10^{-10} \text{ mol dm}^{-3}$

Worked example

A $0.035 \text{ mol dm}^{-3}$ sample of methylamine (CH_3NH_2) has $\text{p}K_b$ value of 3.35 at 298 K. Calculate the pH of methylamine.

Copyright

© 2024 Exam Papers Practice

Answer

Step 1: Calculate the value for K_b using

- $K_b = 10^{-\text{p}K_b}$
 - $K_b = 10^{-3.35}$
 - $K_b = 4.4668 \times 10^{-4}$

Step 2: Substitute values into K_b expression to calculate $[OH^-]$

- $K_b = [OH^-]^2 / [CH_3NH_2]$
- $4.4668 \times 10^{-4} = [OH^-]^2 / 0.035$



- $[\text{OH}^-] = \sqrt{(4.4668 \times 10^{-4} \times 0.035)}$
- $[\text{OH}^-] = 3.9539 \times 10^{-3}$

Step 3: Calculate the pH

- $[\text{H}^+] = K_w \div [\text{OH}^-]$
 - $[\text{H}^+] = (1 \times 10^{-14}) \div 3.9539 \times 10^{-3}$
 - $[\text{H}^+] = 2.5290 \times 10^{-12}$
- $\text{pH} = -\log [\text{H}^+]$
 - $\text{pH} = 2.5290 \times 10^{-12}$
 - $\text{pH} = \mathbf{11.60}$ to 2 decimal places

OR

Step 3: Calculate pOH and therefore pH

- $\text{pOH} = -\log [\text{OH}^-]$
 - $\text{pOH} = -\log 3.9539 \times 10^{-3}$
 - $\text{pOH} = 2.4029$
- $\text{pH} = 14 - \text{pOH}$
 - $\text{pH} = 14 - 2.4029$
 - $\text{pH} = \mathbf{11.60}$ to 2 decimal places