

ACIDS, BASES and PH

Y13 term 2

DEFINITIONS

Acids and bases: The Brønsted-Lowry definition says that **acids** are H^+ (proton) **donors**, and **bases** are H^+ (proton) **acceptors**. The **conjugate base** of an acid is the species that would accept an H^+ ion in the reverse reaction.

E.g. in the reaction of ethanoic acid and ammonia:



EXAM TIP

There are other definitions of acids and bases, e.g. an Arrhenius acid produces H^+ ions in an aqueous solution, and an Arrhenius base produces OH^- ions in an aqueous solution.

18

IONIC EQUATIONS FOR THE REACTIONS OF ACIDS

In ionic equations for the reactions of acids, the acid's anion (we ion) can usually be ignored – it is only the H^+ that reacts, e.g.

Metals form metal ions and hydrogen gas: $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

Carbonates form carbon dioxide and water: $CO_3^{2-}(aq) + 2H^+(aq) \rightarrow CO_2(g) + H_2O(l)$

Metal oxides form metal ions and water: $CaO(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l)$

Hydroxides form water: $OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$

EXAM TIP

Ions that appear (identically) on both sides of a chemical equation are called **spectator ions** and can be left out to simplify the equation

SPECIAL EQUILIBRIUM CONSTANTS, K_a AND K_w

K_a is an equilibrium constant for the dissociation of a weak acid into its ions: $HA \rightleftharpoons H^+ + A^-$

If you assume all H^+ ions come from the acid, HA , then you can say $[H^+] = [A^-]$.

So, by substitution: $K_a = \frac{[H^+][A^-]}{[HA]}$

There is also a logarithmic function: $pK_a = -\log K_a$ and the inverse $K_a = 10^{-pK_a}$

K_w is an equilibrium constant for the dissociation of water, which behaves like an acid. It is called the **ionic product of water**. $H_2O \rightleftharpoons H^+ + OH^-$; $K_w = [H^+][OH^-]$ (H_2O can be left out of the equation as it is effectively constant).

EXAM TIP

The value of K_w changes with temperature because the position of equilibrium moves. You will be given the correct value to use in exam questions.

CALCULATING PH FOR STRONG ACIDS AND BASES

Since a strong acid fully dissociates, $[H^+] = [HA]$ for monobasic acids.

To calculate the pH, find the negative log of the concentration of acid.

Example 1 What is the pH of a 0.04 mol dm^{-3} solution of HCl?

$pH = -\log(0.04) = 1.4$

Example 2 What concentration of H_2SO_4 has a pH of 1.50?

$[H^+] = 10^{-pH} = 10^{-1.50} = 0.032 \text{ mol } dm^{-3}$

Since H_2SO_4 is dibasic, the acid concentration = $0.016 \text{ mol } dm^{-3}$

To calculate the pH of a strong base, use K_w to convert $[OH^-]$ to $[H^+]$

Example 3 What is the pH of a 0.24 mol dm^{-3} solution of NaOH?

Since $K_w = [H^+][OH^-]$ then $[H^+] = \frac{K_w}{[OH^-]} = \frac{1 \times 10^{-14}}{0.24} = 4.2 \times 10^{-14} \text{ mol } dm^{-3}$

$pH = -\log(4.2 \times 10^{-14}) = 13.38$

EXAM TIP

Check the answer makes sense, e.g. for a base, pH must be over 7

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

$K_w = \frac{[H^+][OH^-]}{[H_2O]}$

EXAM TIP

The units of K_a for monobasic acids are always $mol \text{ } dm^{-3}$

CALCULATIONS FOR WEAK ACIDS

A weak acid forms an equilibrium mixture between unionised and ionised forms, and so the equilibrium constant, K_a , must be involved in the calculations.

Example 1: What is the pH of a 0.75 mol dm^{-3} solution of phenol, a weak acid with $K_a = 1.6 \times 10^{-10}$?

$K_a = \frac{[H^+][A^-]}{[HA]}$ so $[H^+] = \sqrt{K_a \times [phenol]} = \sqrt{1.6 \times 10^{-10} \times 0.75} = 1.1 \times 10^{-5} \text{ mol } dm^{-3}$

$pH = -\log(1.1 \times 10^{-5}) = 4.96$

Example 2: What is the K_a of an acid if a 1.5 mol dm^{-3} solution has a pH of 3.42?

$[H^+] = 10^{-pH} = 10^{-3.42} = 3.8 \times 10^{-4} \text{ mol } dm^{-3}$

$K_a = \frac{[H^+]^2}{[HA]} = \frac{(3.8 \times 10^{-4})^2}{1.5} = 9.6 \times 10^{-8} \text{ mol } dm^{-3}$

Example 3: What is the pH of a 0.88 mol dm^{-3} solution of HOCl ($pK_a = 7.50$)?

$K_a = 10^{-pK_a} = 10^{-7.50} = 3.2 \times 10^{-8}$

$[H^+] = \sqrt{K_a \times [HOCl]} = \sqrt{3.2 \times 10^{-8} \times 0.88} = 1.7 \times 10^{-4} \text{ mol } dm^{-3}$

EXAM TIP

pH for weak acids should be less than 7 (but usually not less than 2)

THE PH SCALE

pH is a logarithmic scale that describes the concentration of H^+ ions. To calculate pH, use this formula:

$pH = -\log[H^+]$

If you need to calculate $[H^+]$ use:

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

A pH value of 7 (i.e. a concentration of H^+ ions of $1 \times 10^{-7} \text{ mol } dm^{-3}$) is called **neutral**. pH values less than 7 are acidic (smaller values = more acidic). Those above 7 are basic (higher values = more basic).

The inverse of $\log x$ is 10^x

The negative sign can be confusing – small pH values indicate high concentrations of H^+ ions

LIMITATIONS

These calculations for weak acids involve assumptions and approximations that you need to be aware of.

Assumption 1: $[HA]_{\text{equilibrium}} = [HA]_{\text{initial}}$

The calculations on the left require the concentration of acid at equilibrium, but the initial concentration (before any has reacted) has been used. The approximation is $[HA]_{\text{equilibrium}} = [HA]_{\text{initial}}$

It is possible to calculate the true equilibrium concentration of acid, but this makes the maths harder, and, if the degree of ionisation is very small, $[HA]_{\text{initial}}$ is almost identical to $[HA]_{\text{equilibrium}}$.

If $K_a \gg 1 \times 10^{-4} \text{ mol } dm^{-3}$, enough HA may ionise to cause the approximation to be poor, leading to inaccurate results.

Assumption 2: $[H^+] = [A^-]$

This is true if all H^+ ions come from the same source as the A^- ions (from HA). The only other source of H^+ ions is the solvent, water.

Since K_w is very small ($1 \times 10^{-14} \text{ mol } dm^{-3}$ at 25 °C), the approximation that H_2O does not ionise may cause inaccuracies in very dilute solutions of very weak acids (where $[H^+]$ is small).