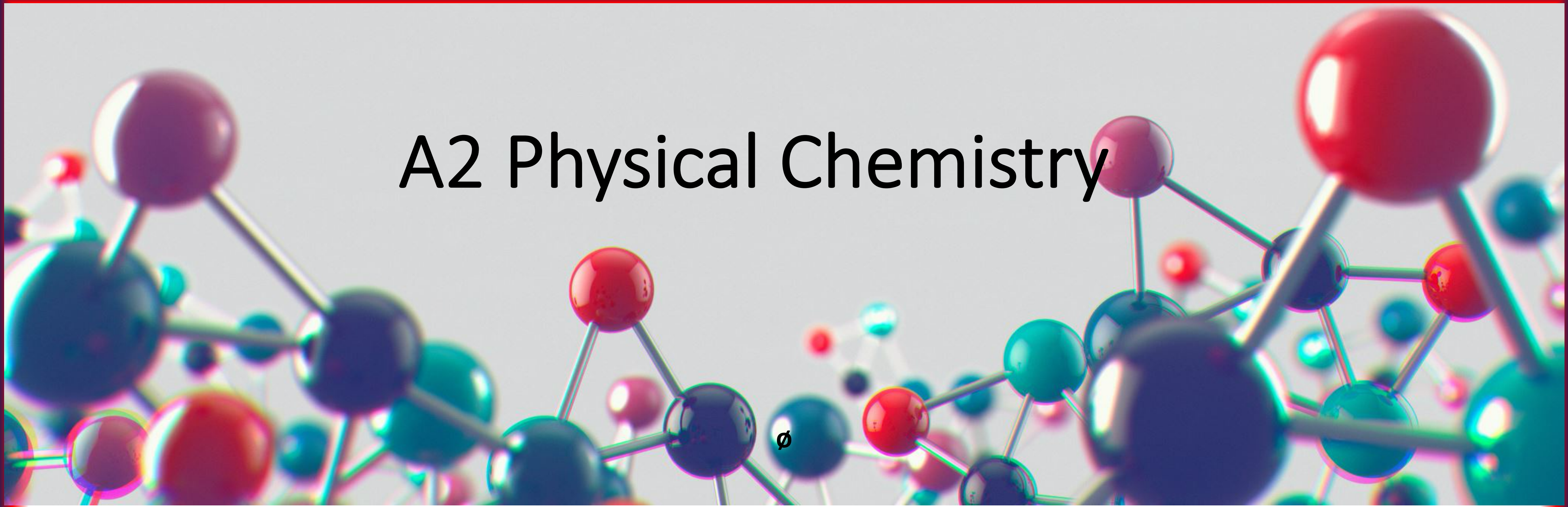
A2 Physical Chemistry



Calculating the pH of Strong Acids

Strong Acids

Strong acids are completely dissociated (broken down) into ions, e.g. hydrochloric and sulphuric acids when added to water.

Both hydrochloric and sulphuric acids are strong acids.

$HCl_{(aq)} \rightarrow H^{+}_{(aq)} + Cl^{-}_{(aq)}$

$H_2SO_{4(aq)} \rightarrow 2H^{+}_{(aq)} + SO_4^{2-}{}_{(aq)}$

Number of Protons Released

Monoprotic acid = **acid that releases one H^+ ion per molecule**

HCl, CH_3COOH, HNO_3

Diprotic acid = **acid that releases two H^+ ions per molecule**

H_2SO_4, HOOC-COOH

Triprotic acid = **acid that releases three H^+ ions per molecule**

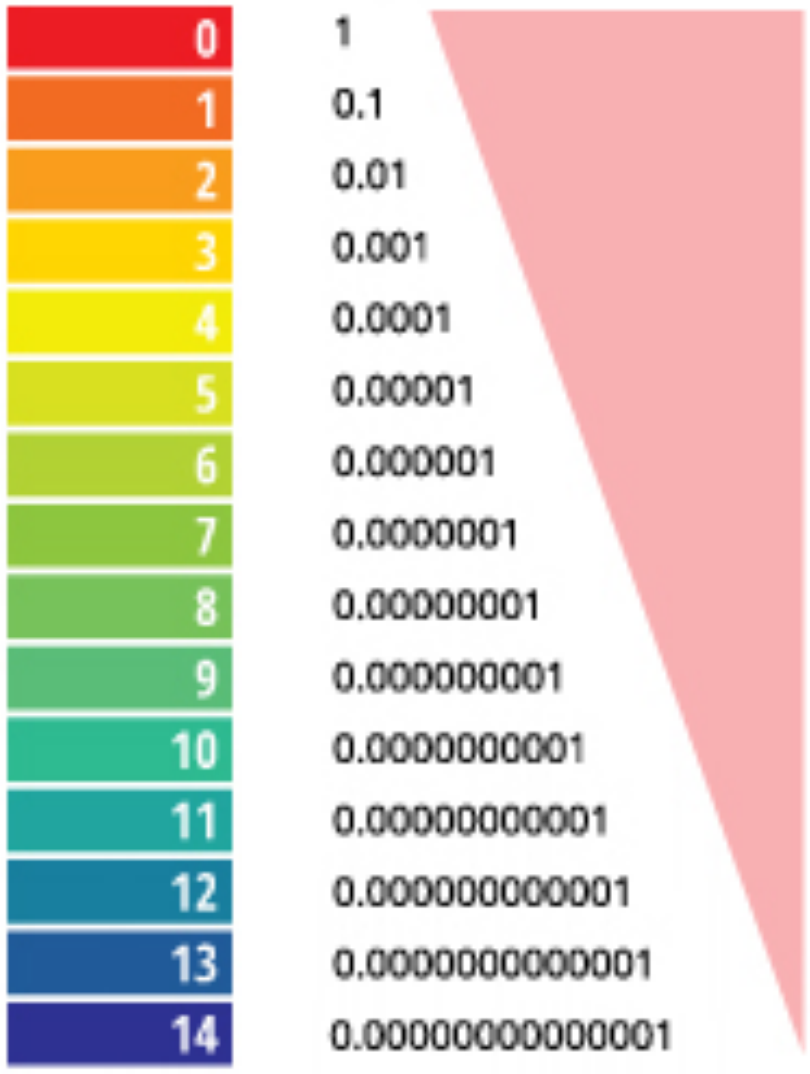
H_3PO_4

The pH Scale

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

The log scale allows a large range of $[H^+]$ to be represented easily.

pH $[H^+]$



Strong Acids - pH calculations 1

What is the pH of 0.20 mol dm^{-3} $HCl_{(aq)}$?

$[H^+]$ = 0.20 mol dm^{-3}

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

$pH = -log_{10}$ (0.2)

pH = 0.70

Strong Acids - pH calculations 2

What is the concentration of a solution of $HNO_{3(aq)}$ with a pH = 1.10?

$[H^+] = 10^{-1.10}$

$[H^+] = 7.94 \times 10^{-2}$ mol dm^{-3}

$[HNO_3\,] = 7.94 \times 10^{-2}$ mol dm^{-3}

Strong Acids - pH calculations 3

What mass of H_3PO_4 is required to make up 250 cm^3 solution of pH 2.35?

$[H^+] = 10^{-0.35}$

$[H^+] = 0.447$ mol dm^{-3}

$[H_3PO_4] = 0.149$ mol dm^{-3}

Mols of H_3PO_4 in 250 cm^3 = 0.149 x 0.250 = 3.72×10^{-2} mol

Mass of H_3PO_4 = 3.72×10^{-2} x 98 = 3.65 g

Strong Acids - pH calculations 4

Calculate the pH of the solution formed when 100 cm^3 of water is added to 50 cm^3 of 0.100 mol dm^{-3} HNO_3.

$[H^+]$ in original solution = 0.100

$$[H^+] \text{ in diluted solution} = 0.100 \times \frac{\text{old volume}}{\text{new volume}}$$

Strong Acids - pH calculations 4

Calculate the pH of the solution formed when 100 cm^3 of water is added to 50 cm^3 of 0.100 mol dm^{-3} HNO_3.

[H^+] in original solution = 0.100

[H^+] in diluted solution = $0.100 \times \frac{50}{150}$ = 0.0333

pH = -log 0.0333

pH = 1.47

Strong Acids - pH calculations 5

Calculate the pH of the solution formed when 250 cm^3 of 0.300 mol dm^{-3} H_2SO_4 is made up to 1000 cm^3 solution with water.

$[H^+]$ in original solution = 2 x 0.300 = 0.600

$$[H^+] \text{ in diluted solution} = 0.600 \times \frac{\text{old volume}}{\text{new volume}}$$

Strong Acids - pH calculations 5

Calculate the pH of the solution formed when 250 cm^3 of 0.300 mol dm^{-3} H_2SO_4 is made up to 1000 cm^3 solution with water.

$[H^+]$ in original solution = 2 x 0.300 = 0.600

$[H^+]$ in diluted solution = $0.600 \times \frac{250}{1000} = 0.150$

pH = -log 0.150

pH = 0.82

Ionic Product of Water

In pure water, a tiny proportion of water molecules are dissociated.

$$H_2O \rightleftharpoons H^+ + OH^-$$

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

$[H_2O]$ is so much larger than $[H^+]$ and $[OH^-]$ it is effectively a constant number.

$$K_c[H_2O] = [H^+][OH^-]$$

Ionic Product of Water

$K_c[H_2O] = [H^+][OH^-]$

This is also a constant

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

pH of Pure Water

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

In pure water $[H^+] = [OH^-]$

$K_w = [H^+]^2$

At 298 K, $K_w = 1.00 \times 10^{-14}\ mol^2\ dm^{-6}$

$1.00 \times 10^{-14} = [H^+]^2$

$[H^+] = 1.00 \times 10^{-7}$

$pH = -\log_{10}(1.00 \times 10^{-7})$

$pH = 7.00$

pH of Pure Water

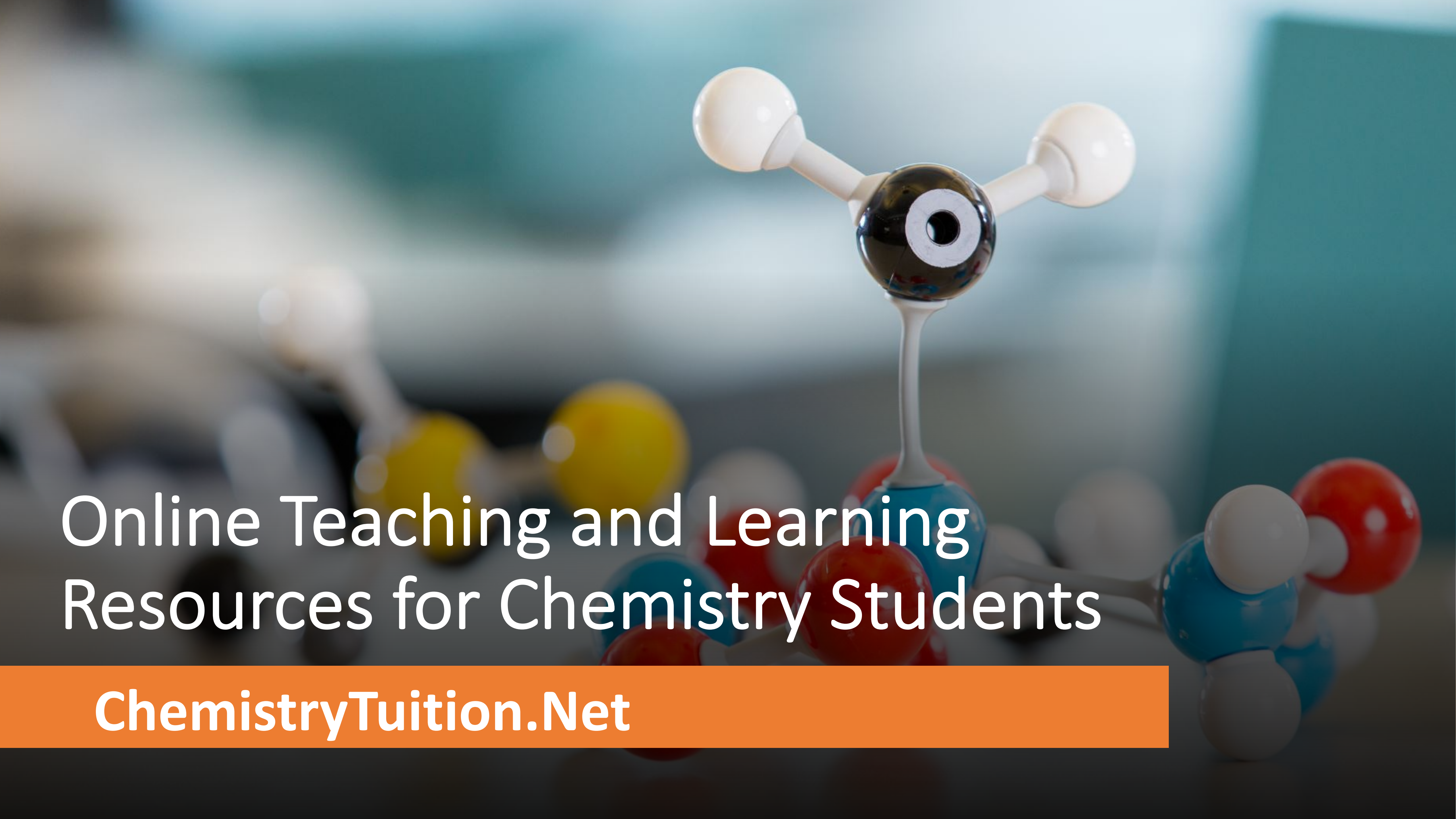
$H_2O \rightleftharpoons H^+ + OH^-$ ΔH = endothermic

As the temperature is increased, the equilibrium shifts towards the products.

Therefore, $[H^+]$ and $[OH^-]$ both increase.

K_w **increases** and pH **decreases**

However the water is still neutral as $[H^+] = [OH^-]$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