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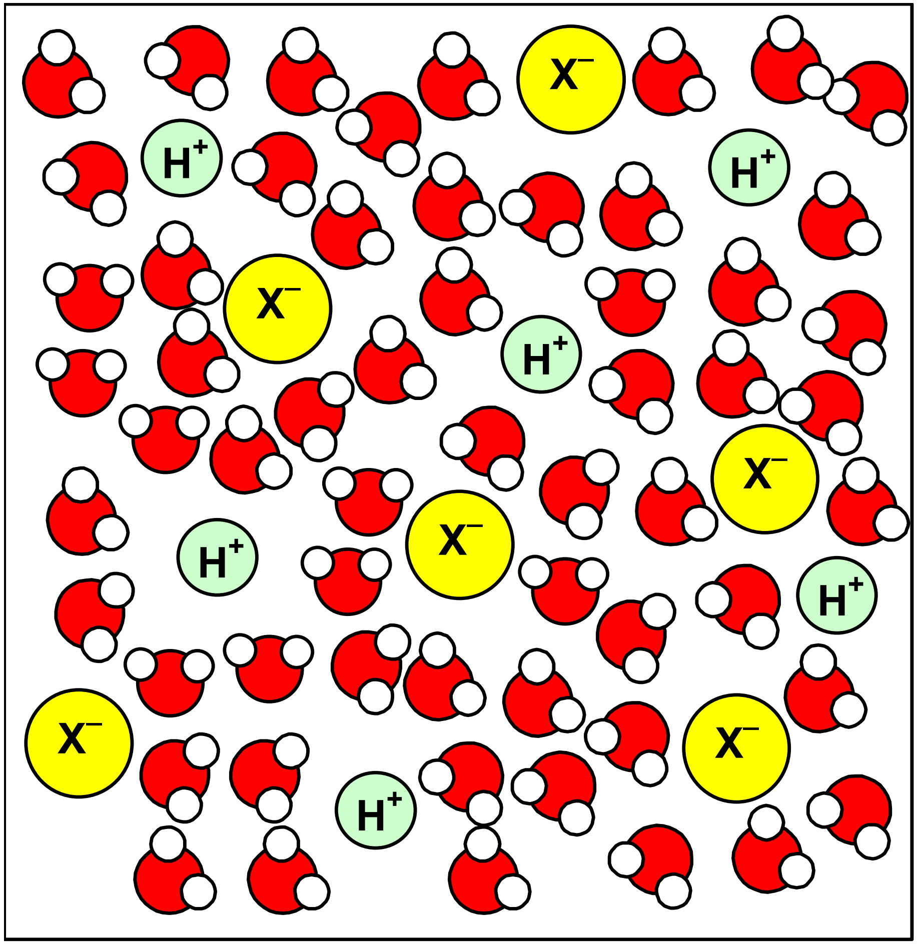
pH of WEAK ACIDS

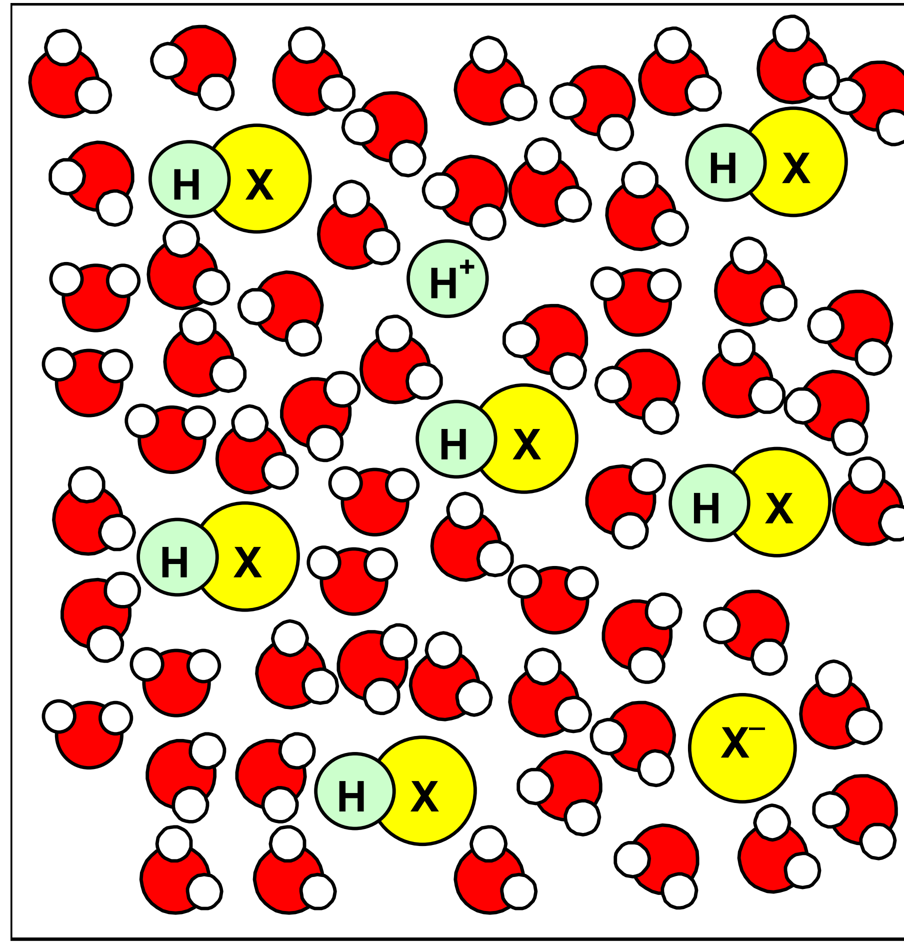
STRONG ACID



$$HX \rightarrow H^+(aq) + X^-(aq)$$

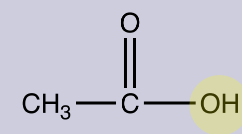
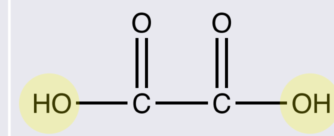
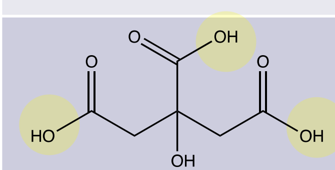
All molecules ionise

WEAK ACID



$$HX \rightleftharpoons H^+(aq) + X^-(aq)$$

Small fraction of molecules ionise

	Strong acids		Weak acids	
Monoprotic	hydrochloric acid nitric acid	HCl HNO_3	ethanoic acid	
Diprotic	sulfuric acid	H_2SO_4	ethanedioic acid	
Triprotic	phosphoric acid*	H_3PO_4	citric acid	

* Not all protons are strong

	Strong bases		Weak bases	
Monobasic	sodium hydroxide	$NaOH$	ammonia	NH_3
	potassium hydroxide	KOH	methylamine	CH_3NH_2
Diprotic	calcium hydroxide	$Ca(OH)_2$	1,2-diaminoethane	$H_2N-CH_2-CH_2-NH_2$
	barium hydroxide	$Ba(OH)_2$		

$$NaOH(s) + aq \rightarrow Na^+(aq) + OH^-(aq)$$

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

$$HA \rightleftharpoons H^+ + A^-$$

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

$$pK_a = -\log K_a$$

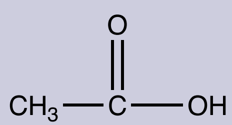
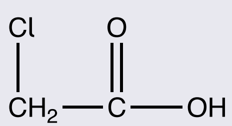
$$K_a = 10^{-pK_a}$$

- K_a = acid dissociation constant
- units are mol dm^{-3}
- the stronger the acid:
  - the bigger K_a
  - the smaller pK_a

$$HA \rightleftharpoons H^+ + A^-$$

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

$$pK_a = -\log K_a$$

Acid		K_a / mol dm^{-3}	pK_a	% dissociation (approx)
Ethanoic acid		1.74×10^{-5}	4.76	0.5%
Chloroethanoic acid		1.38×10^{-3}	2.86	3.7%

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

true for ALL aqueous solutions

In an aqueous solution of a weak acid – IF nothing else added:

$$[H^+] = [A^-]$$

$$K_a = \frac{[H^+]^2}{[HA]}$$

ONLY for aqueous solution opf weak acid with nothing else added

EQUATIONS SUMMARY

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

always give pH to 2dp

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

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

ALL aqueous solutions

$$K_w = [H^+]^2$$

pure water ONLY

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

ALL solutions containing a weak acid

$$K_a = \frac{[H^+]^2}{[HA]}$$

Weak acid in water ONLY

$$pK_a = -\log K_a$$

$$K_a = 10^{-pK_a}$$

pH OF A WEAK ACID

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

ALL solutions containing a weak acid

$$K_a = \frac{[H^+]^2}{[HA]}$$

Weak acid in water ONLY

$pK_a = -\log K_a$

$K_a = 10^{-pK_a}$

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

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

Calculate the pH 0.100 mol dm^{-3} propanoic acid (pK_a = 4.87)

$$K_a = \frac{[H^+]^2}{[HA]} \qquad [H^+]^2 = K_a\,[HA] \qquad [H^+] = \sqrt{K_a\,[HA]}$$

$$[H^+] = \sqrt{10^{-4.87}\ x\ 0.100}$$

$$= 1.16 \times 10^{-3}$$

$pH = -\log [H^+] = -\log 1.16 \times 10^{-3} =$ **2.94**

pH OF A WEAK ACID

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

ALL solutions containing a weak acid

$$K_a = \frac{[H^+]^2}{[HA]}$$

Weak acid in water ONLY

$pK_a = -\log K_a$

$K_a = 10^{-pK_a}$

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

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

Calculate the concentration of a solution of methanoic acid with pH 4.02 (K_a = 1.35 x 10^{-5} mol dm^{-3})

$[H^+] = 10^{-pH} = 10^{-4.02} = 9.55 \times 10^{-5}$

$$K_a = \frac{[H^+]^2}{[HA]} \qquad [HA] = \frac{[H^+]^2}{K_a}$$

$$[HA] = \frac{(9.55 \times 10^{-5})^2}{1.35 \times 10^{-5}}$$

$= 6.76 \times 10^{-4}$ mol dm^{-3}

