



When an acid donates a proton it leaves behind a conjugate base.

Strength

Strong acids completely ionise in water.

Weak acids partially.

Strong bases completely dissociate in water to OH^-

Hydrolysis of Salts.

Acidic anions - HSO_4^- , H_2PO_4^- (anion from polyprotic acid)

Basic anions - CH_3COO^- , HPO_4^{2-} , PO_4^{3-} , SO_4^{2-}
 F^- , HCO_3^- , CO_3^{2-} , ClO^- , HS^- , CN^-
 (anion from weak acid)

Neutral anion: - Cl^- , NO_3^- , Br^- , I^- (anion from strong acid)

Acidic cations - NH_4^+ , Al^{3+} , Fe^{3+} (cation from weak base metal ions)

Basic cations - NONE (we need to know).

Neutral cation - group I + II metals.

a) Na_2SO_4



b) NH_4Cl



Leave out
 any NEUTRAL
 ions
 as spectators.

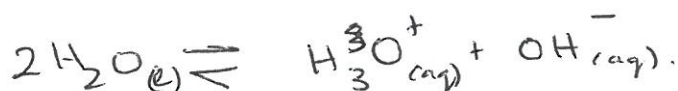
Hydrolysis of Water.

Self-ionisation of water.

Water weak conductor of electricity due to ionisation



or



1 L pure water at 25°C contains $1.00 \times 10^{-7} \text{ H}^+$
and $1.00 \times 10^{-7} \text{ OH}^-$.

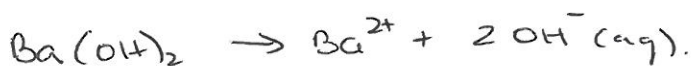
Water is neutral

$$K_w = [\text{H}^+][\text{OH}^-] = 1.00 \times 10^{-14} \text{ at } 25^\circ\text{C}.$$

As temp \uparrow so does K_w (pH changes)
but $[\text{OH}^-] = [\text{H}^+]$ - neutral.

Example.

$[\text{H}^+]$ of a 3.50 mol L^{-1} solⁿ $\text{Ba}(\text{OH})_2$.



$$[\text{OH}^-] = 2 \times [\text{Ba}(\text{OH})_2] = 0.7 \text{ mol L}^{-1}$$

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

$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1 \times 10^{-14}}{0.700}$$

$$= 1.43 \times 10^{-14} \text{ mol L}^{-1}$$

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

↑
conc
known

or pH given
 $= 10^{-\text{pH}}$.

pH < 7 acidic

pH = 7 neutral

pH > 7 basic.

Hydrolysis reactions of weak acids + bases

Acid Reaction

Weak acids - small equilibrium constant.

$$K_a = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

↑
acidity
constant

Base Reaction

Weak bases - small equilibrium constant

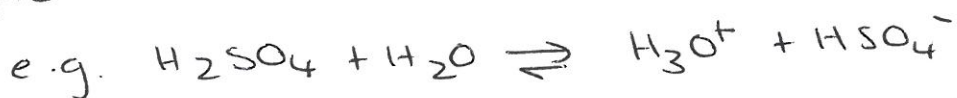
$$K_b$$

↑
basicity
constant.

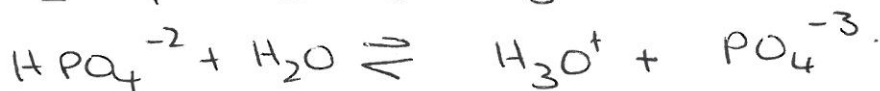
Polyprotic Acids + Bases.

Monoprotic acid - 1 H^+ / molecule of acid

Diprotic acid - 2 ionisable H^+ / molecule of acid.



Triprotic acid - 3 ionisable H^+ / molecule of acid.



Eg. Limiting reagent problem

Calculate the final pH of a solⁿ by mixing 25.0 mL of $0.0750 \text{ mol L}^{-1}$ HCl with 32.5 mL of $0.0675 \text{ mol L}^{-1}$ KOH.

1. Write eqⁿ.



2. Calculate moles.

$$n(\text{HCl}) = c.v = 0.075 \times 0.025 = 1.875 \times 10^{-3} \text{ mol.}$$

$$n(\text{KOH}) = c.v = 0.0675 \times 0.0325 = 2.194 \times 10^{-3} \text{ mol.}$$

3. Ratio. 1:1.

smallest amount HCl.

$$\text{excess } n(\text{KOH}) = n(\text{KOH}) - n(\text{HCl}) \text{ neutralised.}$$

$$= 2.194 \times 10^{-3} - 1.875 \times 10^{-3}$$

$$= 3.188 \times 10^{-4} \text{ mol.}$$

4. Conc. of excess

$$c[\text{KOH}] = \frac{n}{v}$$

$$= \frac{3.188 \times 10^{-4}}{(0.0250 + 0.0325)}$$

$$= 5.54 \times 10^{-3} \text{ mol L}^{-1}$$

Sum of volumes

5. pH calc.

$$[\text{H}^+] = \frac{1.00 \times 10^{-14}}{5.54 \times 10^{-3}}$$

$$= 1.80 \times 10^{-12} \text{ mol L}^{-1}$$

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

$$= 11.7.$$