

# pH and Ionic Equilibria

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#Chemistry

## Acids and Bases

### Bronsted Løwry Theory:

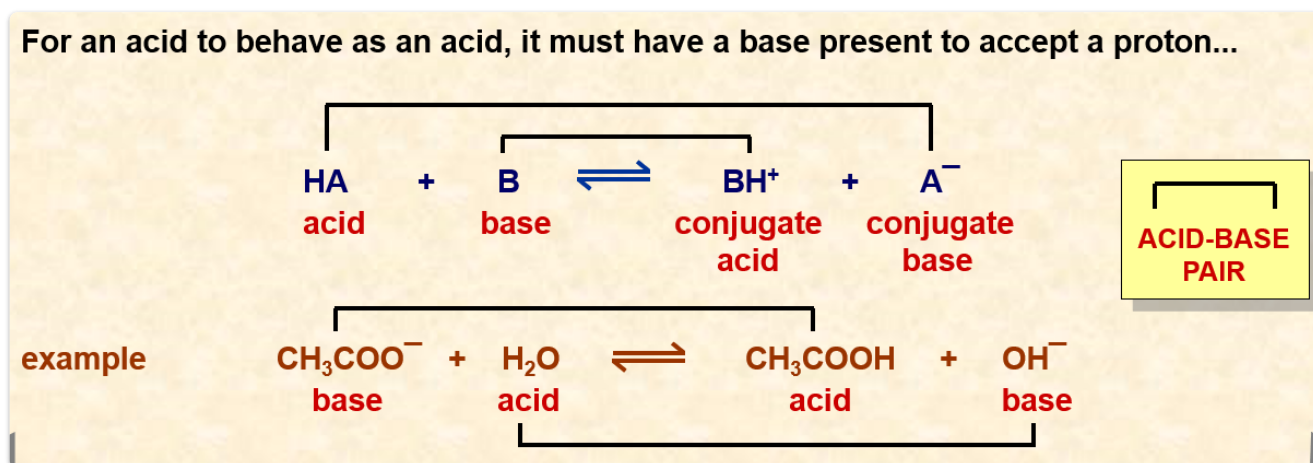
Acids: Proton Donor:  $HCl \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

Base: Proton Acceptor:  $NH_{3(aq)} + H^+_{(aq)} \rightarrow NH_4^+_{(aq)}$

### Conjugate Systems:

Acids are related to bases:  $Acids \rightleftharpoons Proton + Conjugate Base$

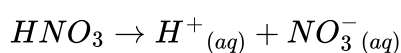
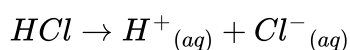
Bases are related to Acids:  $Base + Proton \rightleftharpoons Conjugate Acid$



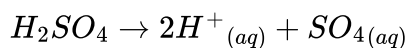
### Strong Acids and Bases:

**Strong Acids** completely dissociate (split up) into ions in aqueous solution.

#### Monoprotic Acids



#### Diprotic Acids



Bases:



**Weak Acids** partially dissociate into ions in aqueous solutions.

💡 E.g. Ethanoic Acid  $CH_3COOH \rightleftharpoons CH_3COO^-_{(aq)} + H^+_{(aq)}$

Conjugate Acid

- *The weaker the acid, the less it dissociates, thus the more the equilibrium lies to the left.*
- The relative strengths of acids can be expressed as  $K_a$  or  $pK_a$  values.

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

$$pK_a = -\log(K_a)$$

The dissociation constant for the weak acid  $HA$  is:

$$K_a = \frac{[H^+_{(aq)}][A^-_{(aq)}]}{[HA_{(aq)}]} \text{ mol dm}^{-3}$$

When a weak base dissolves in water an equilibrium is set up:

*The weaker the base, the less it dissociates, thus the more the equilibrium lies to the left.*

The relative strengths of acids can be expressed as  $K_b$  or  $pK_b$  values.

$$pOH = -\log[OH^-_{(aq)}]$$

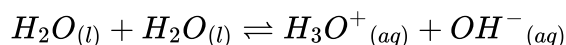
$$pH + pOH = 14$$

[ ] represents the concentration in  $\text{mol dm}^{-3}$ .

## Ionic product of water - $K_w$

Despite being covalent, water conducts electricity to a very small extent.

This is due to slight ionisation:



Applying the equilibrium law to the second equation gives:

$$K_c = \frac{[H^+_{(aq)}][OH^-_{(aq)}]}{[H_2O_{(l)}]}$$

As the dissociation is small, the water concentration is very large compared with the dissociated ions and any changes to its value are insignificant; **its concentration can be**

regarded as constant.

$$K_w = [H^+_{(aq)}][OH^-_{(aq)}] \text{ mol}^2 \text{ dm}^{-6}$$

Because the constant is based on an equilibrium,  $K_w$  *varies with temperature*.

## The Value of $K_w$

The value of  $K_w$  varies with temperature because it is based on an equilibrium.

Temperature / °C	0	20	25	30	60
$K_w / 1 \times 10^{-14} \text{ mol}^2 \text{ dm}^{-6}$	0.11	0.68	1.0	1.47	5.6
$H^+ / \times 10^{-7} \text{ mol dm}^{-3}$	0.33	0.82	1.0	1.27	2.37
pH	7.48	7.08	7	6.92	6.63

This tells us about the equation  $H_2O_{(l)} \rightleftharpoons H^+_{(aq)} + OH^-_{(aq)}$  that:

- $K_w$  gets **larger** as the **temperature increases**
- This means the concentration of  $H^+$  and  $OH^-$  ions gets **greater**
- This means the equilibrium has moved to the **right**
- If the concentration of  $H^+$  **increases** then the **pH decreases**
- **pH** decreases as the temperature **increases**

💡 \*\*Because the equation moves to the right as the temperature goes up, it must be an Endothermic process\*\*

## Relationship between pH and pOH

Because  $H^+$  and  $OH^-$  ions are produced in equal amounts when water dissociates, their concentrations will be the same.

[H <sup>+</sup> ]	10 <sup>0</sup>	10 <sup>-1</sup>	10 <sup>-2</sup>	10 <sup>-3</sup>	10 <sup>-4</sup>	10 <sup>-5</sup>	10 <sup>-6</sup>	10 <sup>-7</sup>	10 <sup>-8</sup>	10 <sup>-9</sup>	10 <sup>-10</sup>	10 <sup>-11</sup>	10 <sup>-12</sup>	10 <sup>-13</sup>	10 <sup>-14</sup>
OH <sup>-</sup>	10 <sup>-14</sup>	10 <sup>-13</sup>	10 <sup>-12</sup>	10 <sup>-11</sup>	10 <sup>-10</sup>	10 <sup>-9</sup>	10 <sup>-8</sup>	10 <sup>-7</sup>	10 <sup>-6</sup>	10 <sup>-5</sup>	10 <sup>-4</sup>	10 <sup>-3</sup>	10 <sup>-2</sup>	10 <sup>-1</sup>	10 <sup>0</sup>
pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14
pOH	14	13	12	11	10	9	8	7	6	5	4	3	2	1	0
STRONGLY ACIDIC				WEAKLY ACIDIC			NEUTRAL		WEAKLY ALKALINE			STRONGLY ALKALINE			

## Calculating the pH of acids and alkalis

💡 Calculating the \$pH\$ of 0.02M HCl:

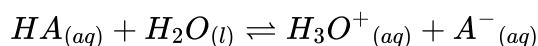


$$[H^+] = 0.02M = 2 \times 10^{-2} \text{mol dm}^{-3}$$

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## The dissociation constant for a weak acid - $K_a$

A weak monobasic acid (HA) dissociates in water thus.

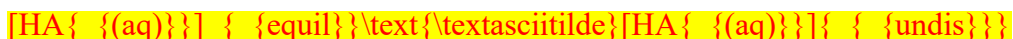


Applying the equilibrium law, we get:

$$K_c = \frac{[H_3O^+_{(aq)}][A^-_{(aq)}]}{[HA_{(aq)}][H_2O_{(aq)}]}$$

### Assumptions:

- For a weak acid, there is little dissociation.

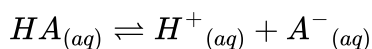


This assumption becomes less valid for stronger weak acids where there is more dissociation.

In dilute solution, the concentration of water is large compared with the dissociated ions and any changes to its value are insignificant; its concentration can be regarded as 'constant'.

$H_2O$  is constant

A weak acid dissociates as follows:



$$K_a = \frac{[H^+_{(aq)}][A^-_{(aq)}]}{[HA_{(aq)}]} \text{ mol dm}^{-3}$$