2. Acids and Basesy

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考点一：酸碱定义

2.1 classify substances as acids, bases, amphoteric or salts, based on their characteristics, name, and formula

2.4 describe various acid-base definitions up to the Brønsted-Lowry definition including Arrhenius, Brønsted-Lowry and Lewis

2.6 compare strong and weak acids and bases using the concept of equilibrium

1. Acid and base

1.1 Arrhenius

Acid ionizes to produce H^+ ions. **Base** dissociates to produce OH^- ions.

$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$ $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$

1.2 Brønsted-Lowry

Acid is a proton donor. eg, ammonium ions (NH_4^+).

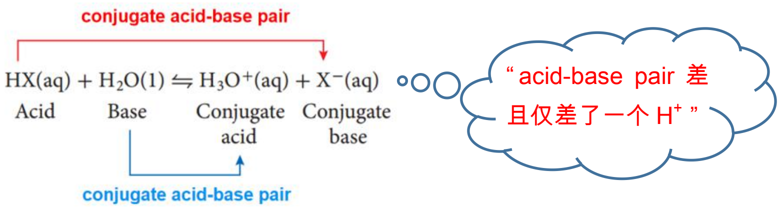
Base is a proton acceptor. eg, hydroxide ions (OH^-).

Brønsted–Lowry conjugate acid-base pairs

• A **conjugate acid** is the species produced when a base accepts a hydrogen ion(接受到了 H^+后).

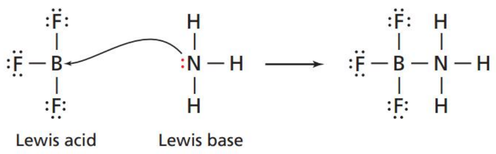
• A **conjugate base** is the species produced when an acid donates a hydrogen ion(失去了 H^+后).

• A **conjugate acid-base pair** consists of two substances related to each other by donating and accepting a single hydrogen ion.



1.3 Lewis

A **Lewis acid** is an electron-pair acceptor and a **Lewis base** is an electron pair donor.



2. Strong acid and weak acid

--A **strong acid** is completely dissociates to ions when in solution, with pH 0-1.

$$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$$

HCl（盐酸）、H_2SO_4（硫酸）、HNO_3（硝酸）、HBr（氢溴酸）、HI（氢碘酸）

--A **weak acid** only slightly dissociates when in solution, with pH between 3 and 7.

$$CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + H^+(aq)$$

--The same definitions are true for **strong** and **weak bases**. 弱碱：$NH_3{\cdot}H_2O$

3. Amphoteric

A substances that can act as acids or bases are called **amphoteric**（两性的）.

$$HF(aq) + H_2O(l) \leftrightarrow H_3O^+(aq) + F^-(aq); \quad NH_3(aq) + H_2O(l) \leftrightarrow NH_4^+(aq) + OH^-(aq)$$

考点二：pH 计算

2.7 calculate the pH of a strong or weak acid or base given its concentration, and vice versa

1. pH of a strong acid&base

求 pH 就是求[H⁺]

$$pH = -\log[H^+];\quad pOH = -\log[OH^-];$$

$$pH + pOH = 14$$

2. pH of a weak acid&base

$$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$$

$$K_a = \frac{[H^+(aq)][A^-(aq)]}{[HA(aq)]} = \frac{[H^+]^2}{[HA]}$$

$$[H^+(aq)] = \sqrt{(K_a \times [acid])}$$

$$[OH^-(aq)] = \sqrt{(K_a \times [base])}$$

3, **Ka & pKa**

$$pK_a = -\log_{10}K_a$$

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

The **larger** the value of Ka the **stronger** the acid. The **smaller** the value of pKa the **stronger** the acid.

Ka 越大越好， pKa 越小越好！

3. K_w & pK_w

Water slightly dissociates to form **hydroxide** and **hydrogen ions** as an equilibrium with its own **equilibrium constant**, K_w

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

At **25°C**, **room temperature**, K_w has a constant value of **1.0×10^{-14}**. However, as temperature changes, this value changes

2, Use K_w to Find the pH of a Strong Base

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

Rearrange Kw to find $[H^+]$: $[H^+] = \dfrac{Kw}{[OH^-]}$

Use $Kw = 10^{-14}$ and the concentration of the base to find $[H^+]$.

Finally, calculate pH using: $pH = -\log_{10}[H^+]$

$$pK_w = -\log_{10}K_w$$

$$K_w = 10^{-pK_w}$$

The higher temperature, the higher K_a, the smaller pK_w

K_a，和 pK_w 的变化趋势总算相反!!

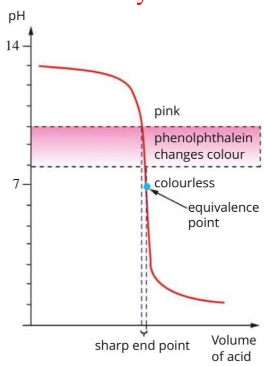
考点三：Neutralization and Titration

2.2 describe how neutralization involves tempering the effects chemical characteristics of an acid with a base or vice versa
2.3 demonstrate a knowledge of WHMIS standards by selecting and applying proper techniques for handling and disposing of lab materials (e.g., use appropriate techniques for handling and disposing of acids and bases)
2.5 predict products of acid-base reactions
2.8 describe the interactions between H^+ ions and OH^- ions using Le Châtelier's principle
2.9 determine the concentration of an acid or base solution using stoichiometry
2.10 explain how acid-base indicators function

1. End point&equivalence point

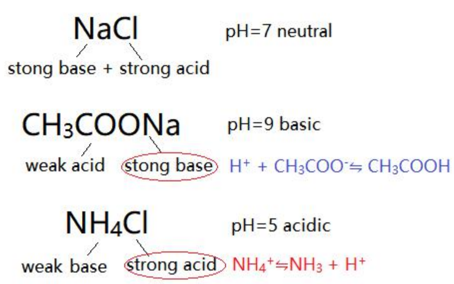
The **end point (变色点)** is the point during the titration at which the indicator changes colour

The **equivalence point(完全反应点)** is the point when the amount in moles of the reactant added from a burette is just enough to react exactly with all of the measured amount of chemical in the flask



"end point"不一定是
"equivalence point"！

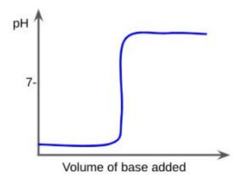
2, Salt



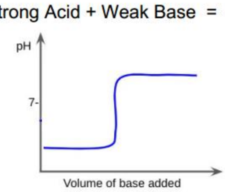
3. **Titration Curves**

A pH titration curve shows how pH of a solution **changes** during an **acid-base reaction**. When an acid and base react, a **neutralisation point** is reached, identified as a large **vertical section** of the graph the pH titration curve.

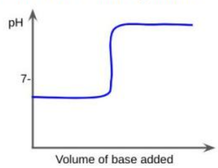
Strong Acid + Strong Base = pH 7



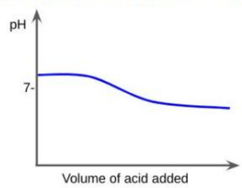
- Strong Acid + Weak Base = < pH 7 (more acidic)



- Weak Acid + Strong Base = > pH 7 (more basic)



- Weak Acid + Weak Base = normally pH 7 but hard to determine

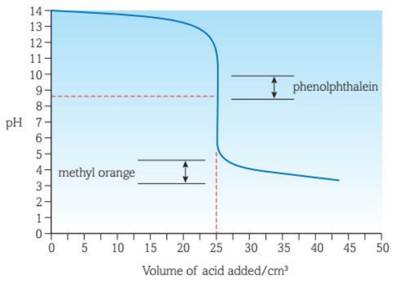


4, Indicators

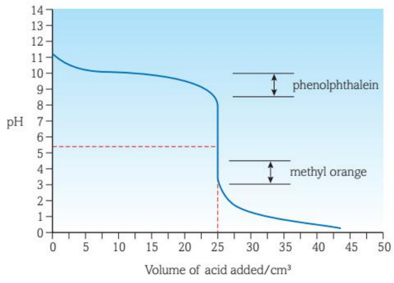
Specific chemical indicators have to be used for specific reactions as they can only indicate a pH change within a certain range

The two most common indicators used at A-Level are methyl orange and **phenolphthalein** :

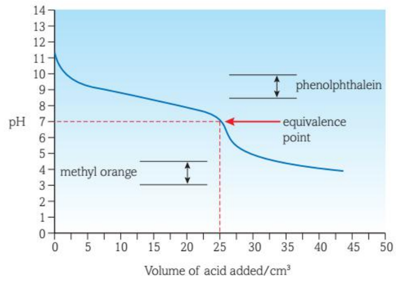
Indicator	pH at colour change	Colour in acid	Colour in base
Methyl orange	3-5	Red	Yellow
Phenolphthalein	8-10	Colourless	Pink
Litmus	5-8	Red	Blue



the pH range of phenolphthalein falls within the steep section



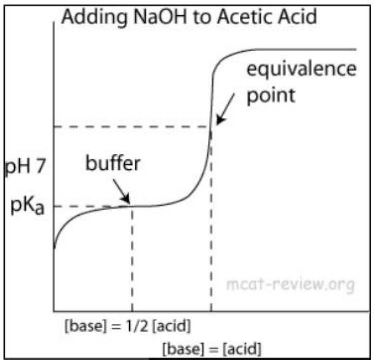
pH range of methyl orange falls within the steep section



no prominent steep section

5. **Calculating Ka from Titration Curve**

The vertical region of a titration curve is the **equivalence point**. At half-equivalence point (when half the volume of acid/base at the equivalence point is added) the pH is equal to pKa, by definition.



Therefore, by reading the pH at half the equivalence point, Ka can be easily calculated from pKa.

$$[CH_3COOH(aq)] = [CH_3COO^-(aq)]$$

$$\text{So: } [H^+(aq)] = \frac{K_a \times [CH_3COOH(aq)]}{[CH_3COO^-(aq)]} = K_a$$

$$\text{Hence, at this point:} \quad [H^+(aq)] = K_a$$

$$pH = pK_a$$

考点四：Buffer

2.11 explain how buffers function

Calculating the pH of a buffer solution

Example: A buffer solution contains 0.35 mol dm^{-3} methanoic acid and 0.67 mol dm^{-3} sodium methanoate. For methanoic acid, $K_a = 1.6 \times 10^{-4}$ mol dm^{-3}. Find the pH of this buffer.

We assume that the sodium methanoate completely dissociates so that the equilibrium concentration of $HCOO^-$ is the same as the initial concentration of $HCOO^-Na^+$. Similarly, HCOOH only slightly dissociates so we assume that the equilibrium concentration is equal to the initial concentration .

1) **First find the expression for Ka for methanoic acid**

$$HCOOH \rightleftharpoons H^+ + HCOO^- \qquad Ka = \frac{[H^+] \times [HCOO^-]}{[HCOOH]}$$

2) Rearrange the expression to find $[H^+]$

$$[H^+] = \frac{Ka \times [HCOOH]}{[HCOO^-]}$$

$[H^+] = 1.6 \times 10^{-4} \times (0.35/0.67) = 8.4 \times 10^{-5}$

$$[H^+] = K_a \times \frac{[acid]}{[salt]}$$

记忆：
“Ka(acid)/b(base)”

3). Convert $[H^+]$ to pH

$pH = -\log_{10} (8.4 \times 10^{-5}) = 4.08$