Unit 8

Acids & Bases

8.1 Introduction to Acids & Bases
8.2 pH & pOH of Strong Acids & Bases

- Autoionization of Water
- pH of Strong Acids
- pOH & pH of Strong Bases

Terminology

Arrhenius *(what you learned last year)*

- Acid: substances that dissolve in water to form H^+
- Base: substances that dissolve in water to form OH^-

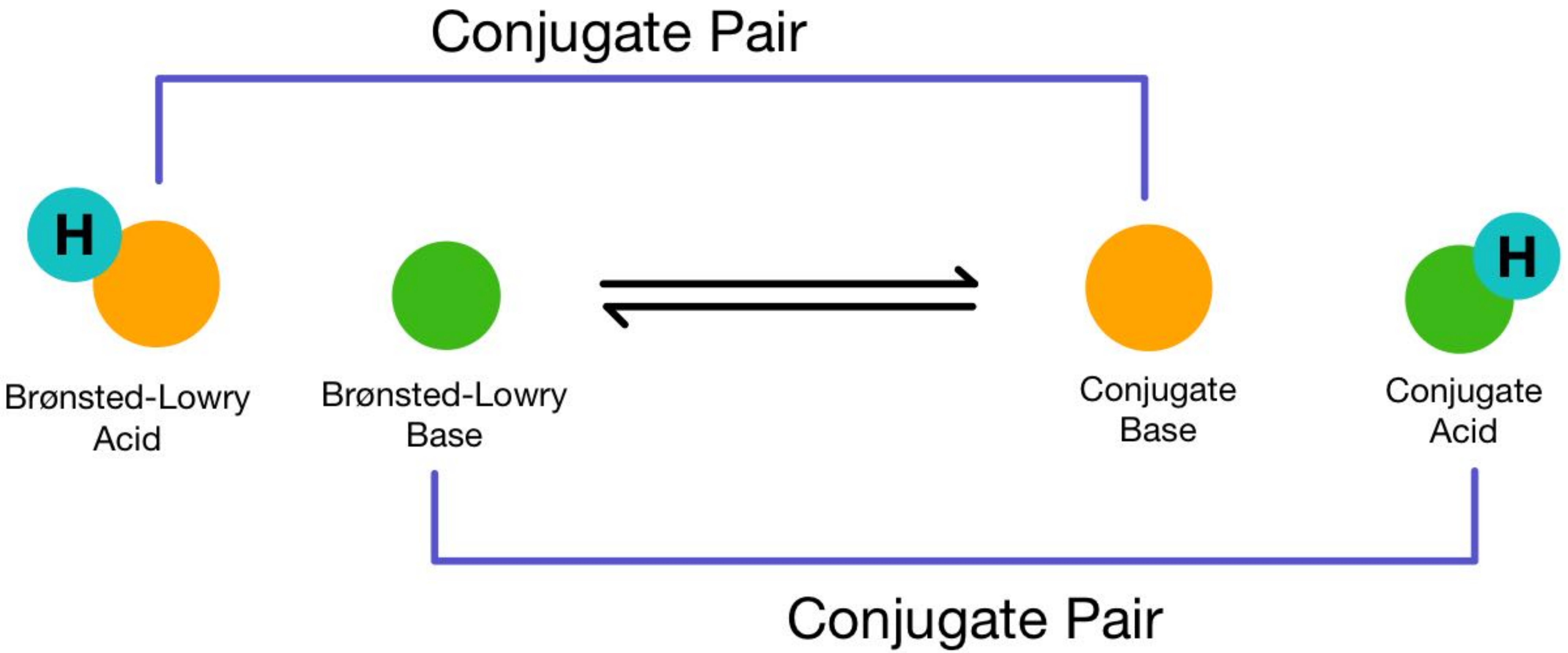
Brønsted-Lowry *(transfer of H^+ during an acid-base reaction)*

- Acid: donates H^+
- Base: accepts H^+

Acid Base

$$HCl + NH_3 \Leftrightarrow NH_4^+ + Cl^-$$

Conjugate Pairs



Acid Base CA CB

$HCl + NH_3 \Leftrightarrow NH_4^+ + Cl^-$

Example: Conjugate Pairs

Label the conjugate acid-base pairs in the reaction.

In a conjugate pair, the acid will always have one more proton than its conjugate base - making it easier to predict formulas

$$CH_3COOH_{(aq)} \;\; H_2O_{(l)} \rightleftarrows CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$$

Autoionization of Water

$$2\ H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

In pure water at 25°C,

$[H_3O^+] = 1.0 \times 10^{-7}$ and $[OH^-] = 1.0 \times 10^{-7}$.

As $[H_3O^+] = [OH^-]$ the solution is neutral.

$$K_w = [H_3O^+]\ [OH^-]$$

$$K_w = (1.0 \times 10^{-7})\ (1.0 \times 10^{-7})$$

$$K_w = 1.0 \times 10^{-14} \quad \text{(at 25 °C)}$$

Autoionization of Water

$$2\,H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + OH^-_{(aq)}$$

$$K_w = [H_3O^+]\,[OH^-] = 1.0 \times 10^{-14} \quad (\text{at } 25\ ^{o}\text{C})$$

- This is a constant.
- If you know the concentration of $[H_3O^+]$ you can calculate the concentration of $[OH^-]$ or vise versa.
- Memorize this constant! It’s very useful.

Example: Autoionization of Water

- The concentration of OH^- in a solution is found to be 1.4×10^{-11} M at 25°C. Find $[H_3O^+]$

H^+ and H_3O^+

- Equations for acids can be written as:

$$HA(aq) + H_2O(l) \rightarrow H_3O^+(aq) + A^-(aq)$$

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

Hydrogen ion, H^+, and hydronium ion, H_3O^+, are used interchangeably for the aqueous ion of hydrogen.

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

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

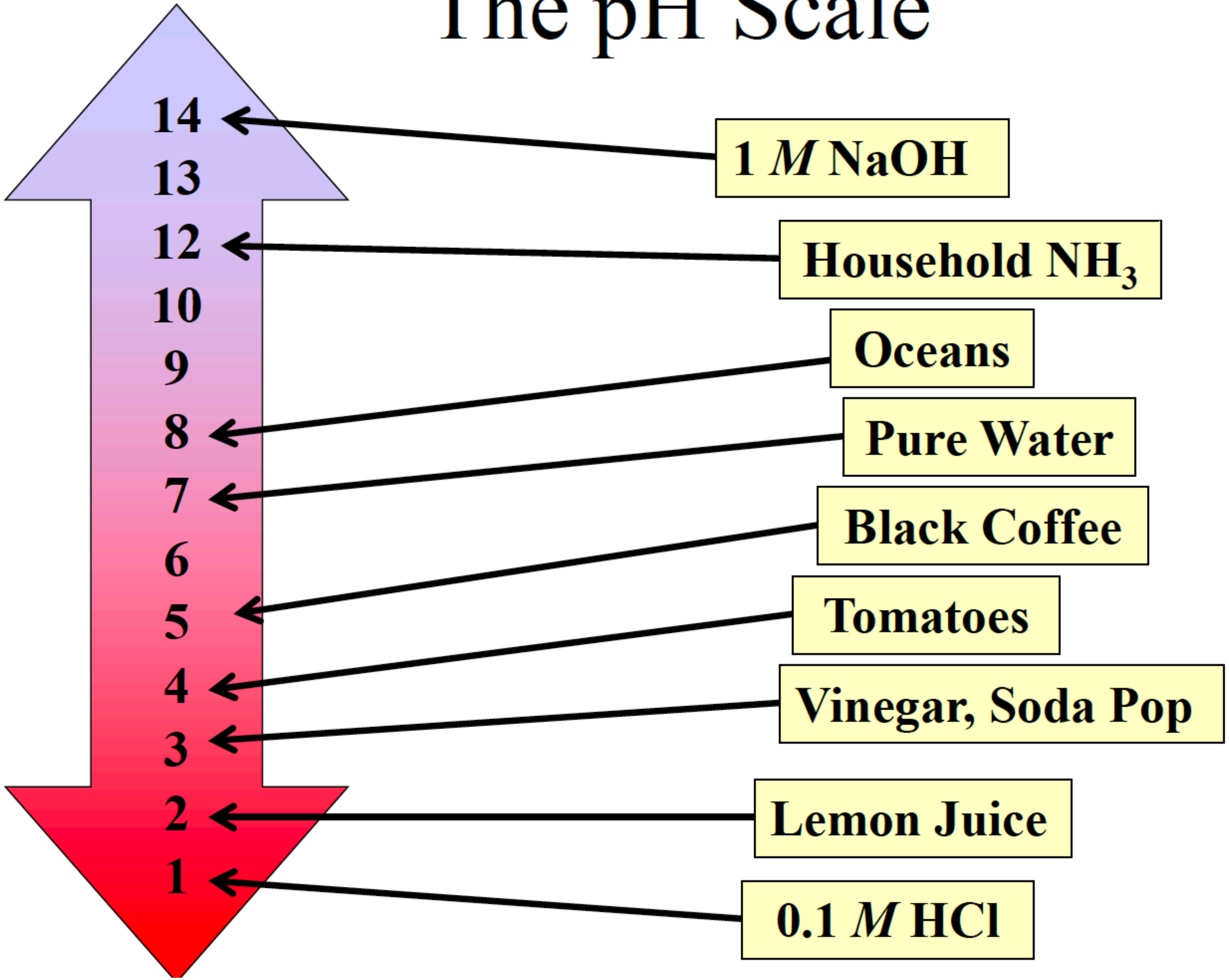
$[OH^-] = 10^{-pOH}$ M

- Expresses the negative power of 10 for the concentration of H^+
- A solution with $[H_3O^+] = 10^{-4}$ has a pH of 4

Find the pH of pure water at 25°C, remembering that $K_w = 1.0 \times 10^{-14}$ at 25°C.

If a solution has a pH of 5.63, find the $[H^+]$.

The pH Scale


14
13
12
10
9
8
7
6
5
4
3
2
1
1 *M* NaOH
Household NH_3
Oceans
Pure Water
Black Coffee
Tomatoes
Vinegar, Soda Pop
Lemon Juice
0.1 *M* HCl

The ‘Big Six’ Strong Acids

- These are the only acids that completely ionize (dissociate) in aqueous solution to produce hydronium ions, H_3O^+.

$HClO_4$	HBr
HI	HNO_3
HCl	H_2SO_4

Example: Calculating pH for a Strong Acid Solution

- Find the pH of a 2.5 *M* HCl solution.

For strong acids, we assume 100% ionization.

- Find the pH of a 0.05 *M* H_2SO_4 solution.

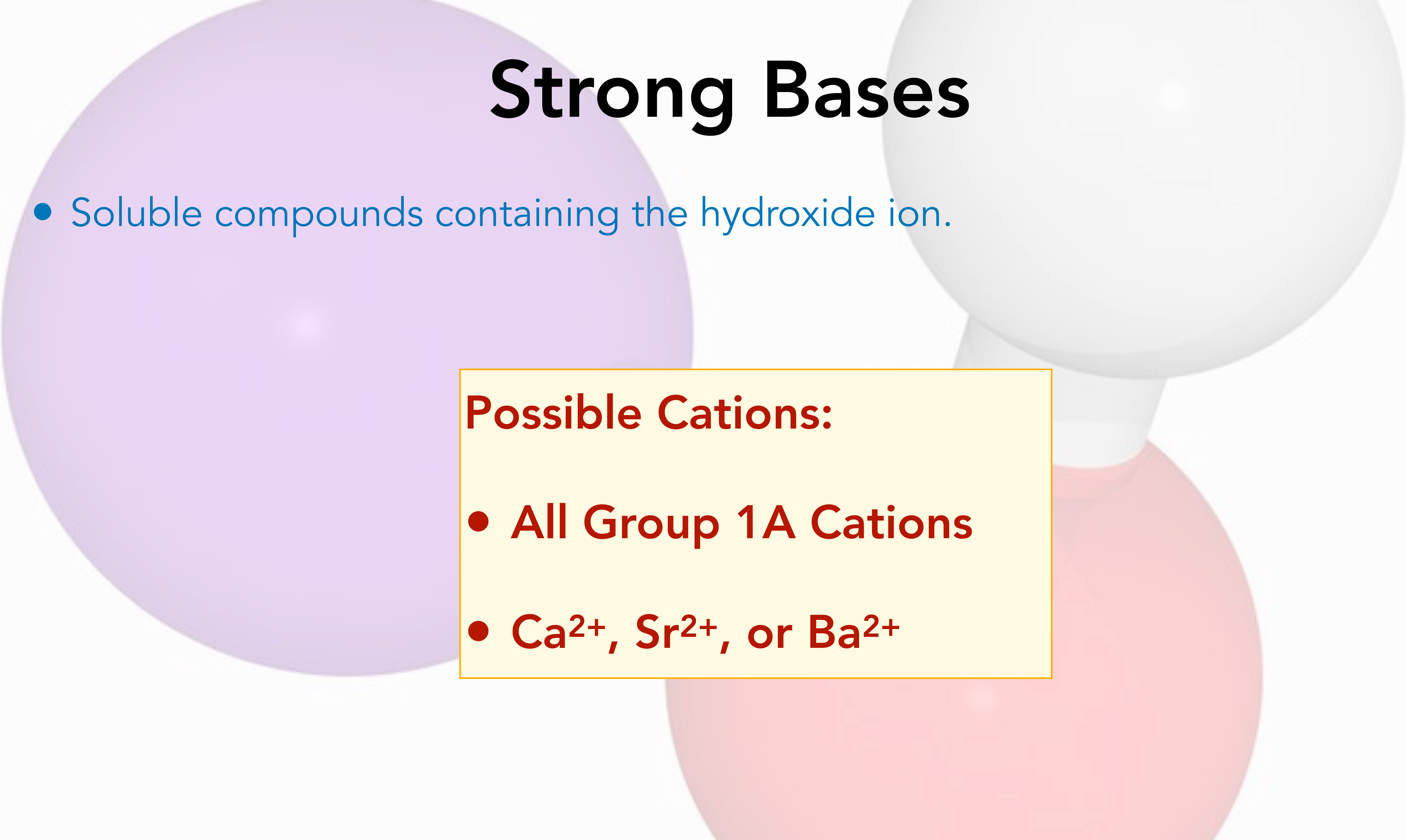
Example: Calculating pH for a Strong Acid Solution

Find the pH of a solution that was created by mixing 500.0 mL of 0.15 M HBr with 250.0 mL of 0.20 M HNO_3. Assume the volumes are additive.

Step 1: Find the concentration of H_3O^+.

Step 2: Find the final pH.

Strong Bases

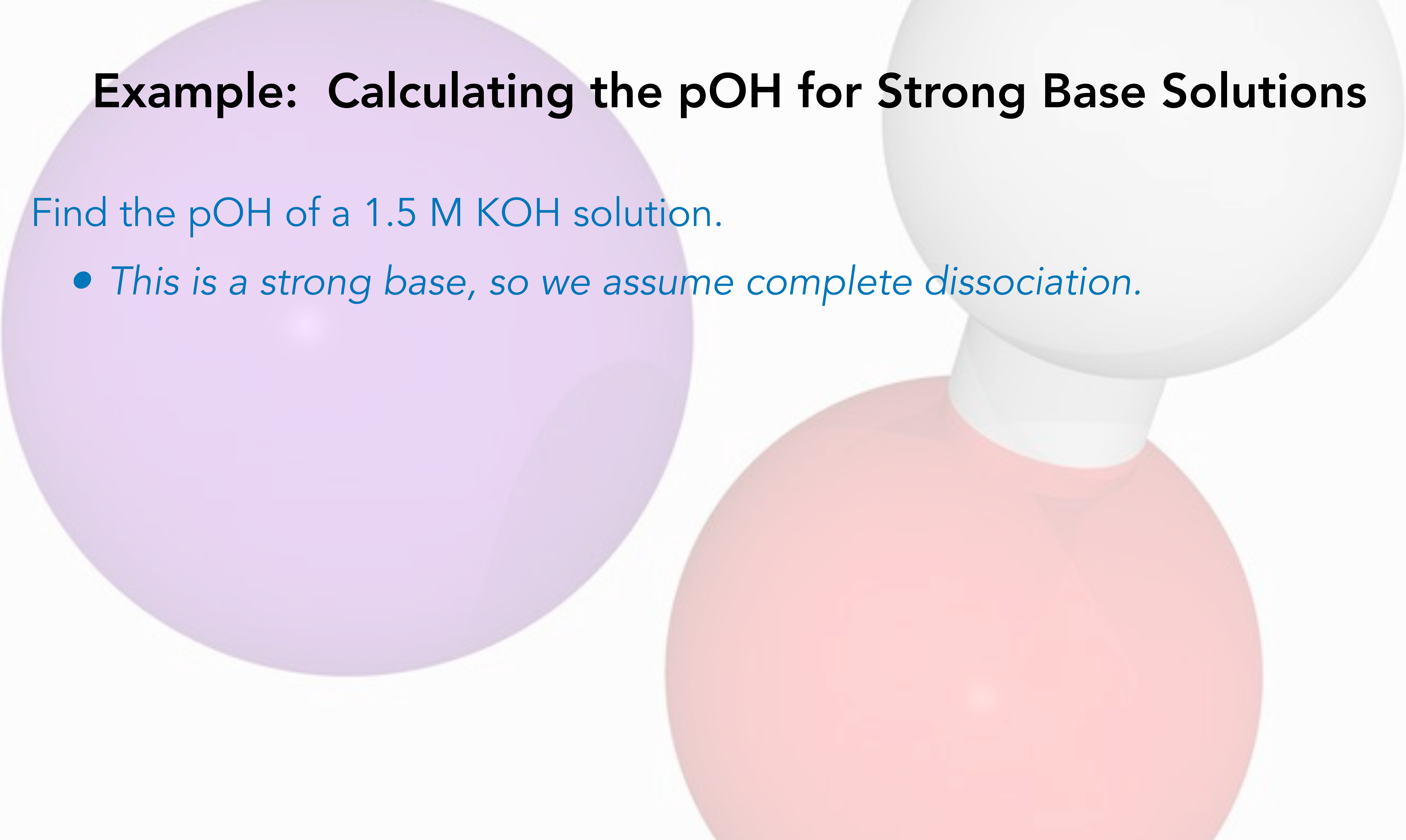


- Soluble compounds containing the hydroxide ion.

Possible Cations:

- **All Group 1A Cations**
- **Ca^{2+}, Sr^{2+}, or Ba^{2+}**

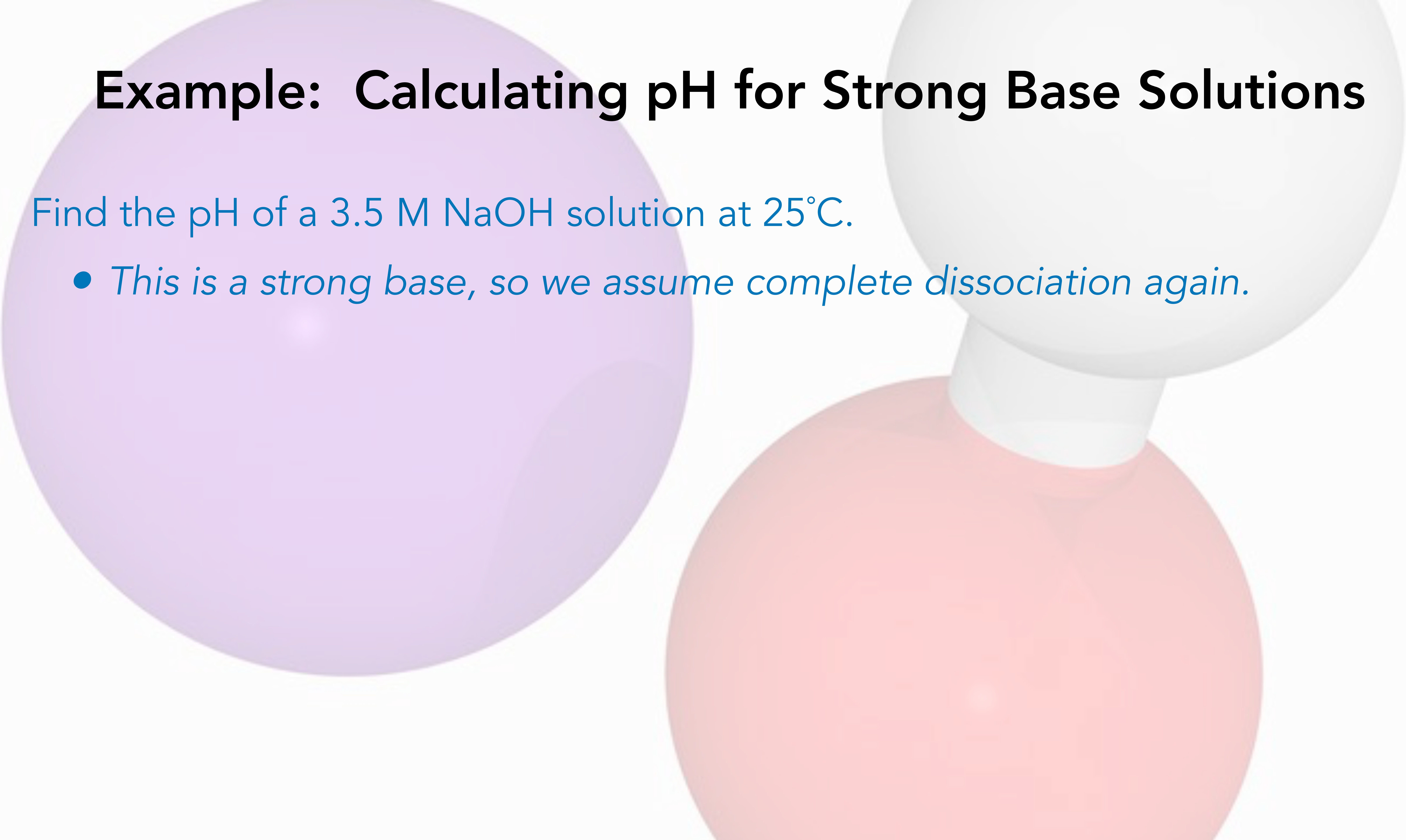
Example: Calculating the pOH for Strong Base Solutions



Find the pOH of a 1.5 M KOH solution.

- *This is a strong base, so we assume complete dissociation.*

Example: Calculating pH for Strong Base Solutions



Find the pH of a 3.5 M NaOH solution at 25°C.

- *This is a strong base, so we assume complete dissociation again.*

K_w at Temperatures Other than 25°C

$$2\,H_2O + \text{heat} \rightleftharpoons H_3O^+ + OH^-$$

$$K_w = [H_3O^+][OH^-]$$

- The auto-dissociation of water is an endothermic process.
- When temperature increases, the equilibrium shifts to the right, so $[H_3O^+]$, $[OH^-]$ and K_w increase.
- When the temperature decreases, the equilibrium shifts to the left so $[H_3O^+]$, $[OH^-]$ and K_w decrease.

8.3 Weak Acid & Base Equilibria

- K_a and K_b
- pH of Weak Acids & Bases
- Polyprotic Acids

Acid Ionization Constant (K_a)

$$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + NH_{3(aq)}$$

$$K_a = \frac{[H_3O^+][NH_3]}{[NH_4^+]}$$

Acid	**K_a**
HCl	2×10^6
HNO_3	30
HSO_4^-	0.011
$C_9H_8O_4$	6.6×10^{-4}
CH_3COOH	1.8×10^{-5}
HCN	4.9×10^{-10}

The stronger the acid, the larger the K_a

Base Ionization Constant (K_b)

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons OH^-_{(aq)} + NH_4^+{}_{(aq)}$$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

Base	**K_b**
NaOH	Very Large Number
NH_3	1.8×10^{-5}
ClO^-	3.3×10^{-7}
H_2NOH	1.1×10^{-8}
C_5H_5N	1.7×10^{-9}

The stronger the base, the larger the K_b

pK_a and pK_b Values are Often Given

$$\mathrm{p}K_\mathrm{a} = -\log_{10}(K_\mathrm{a})$$

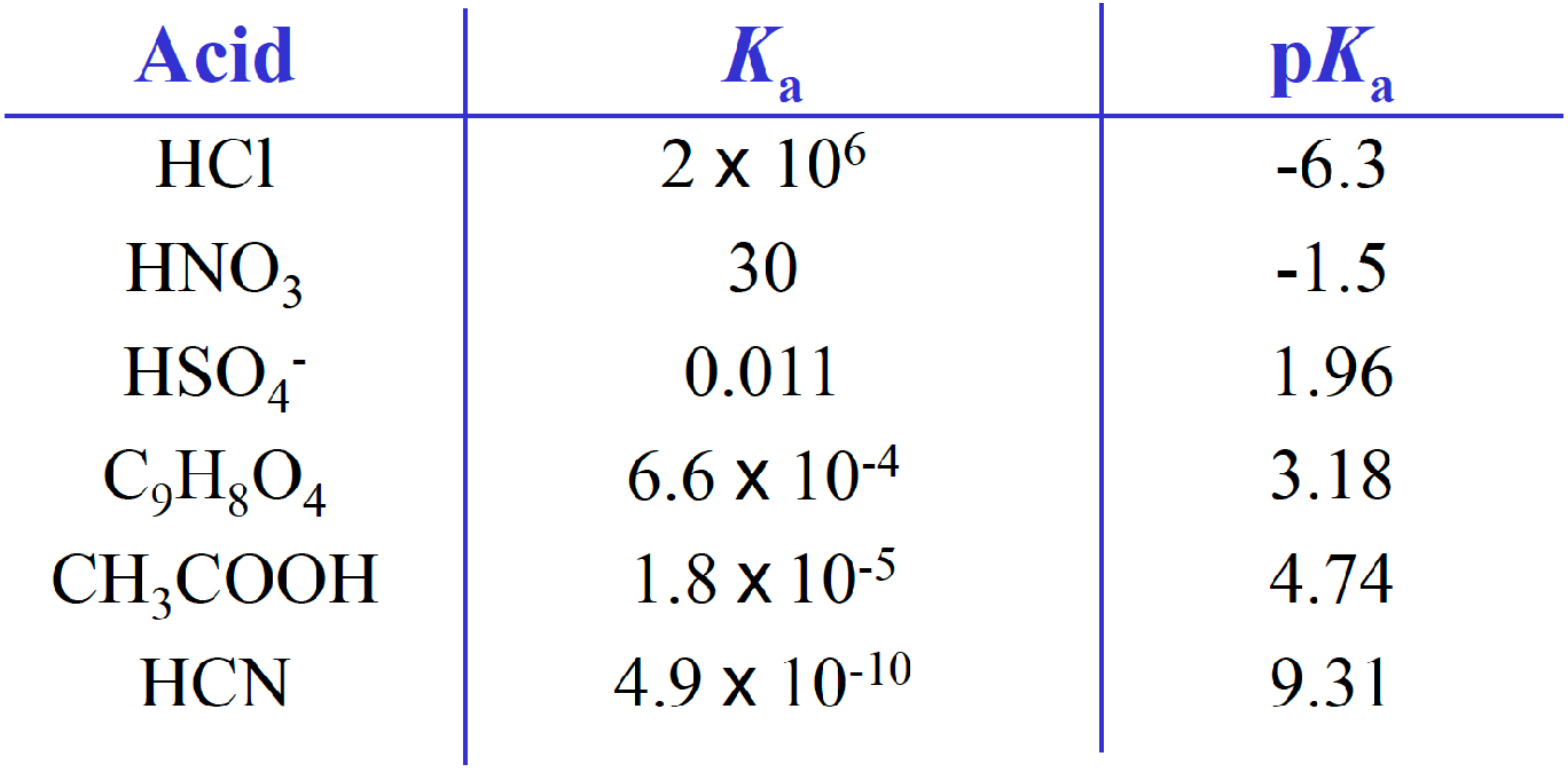
$$K_\mathrm{a} = 10^{-\mathrm{p}K_\mathrm{a}}$$

$$\mathrm{p}K_\mathrm{b} = -\log_{10}(K_\mathrm{b})$$

$$K_\mathrm{b} = 10^{-\mathrm{p}K_\mathrm{b}}$$

pK_a for HCN is 9.31 at 25°C. Find K_a at 25°C.

pK_a



Acid	K_a	pK_a
HCl	2×10^{6}	-6.3
HNO_3	30	-1.5
HSO_4^-	0.011	1.96
$C_9H_8O_4$	6.6×10^{-4}	3.18
CH_3COOH	1.8×10^{-5}	4.74
HCN	4.9×10^{-10}	9.31

The stronger the acid, the smaller the pK_a

pK_b

Base	**K_b**	**pK_b**
NaOH	Very Large Number	Negative Number
NH_3	1.8×10^{-5}	4.74
ClO^-	3.3×10^{-7}	6.48
H_2NOH	1.1×10^{-8}	7.96
C_5H_5N	1.7×10^{-9}	8.77

The stronger the base, the smaller the pK_b

Relationship Between K_a and K_b

$$K_w = K_a \times K_b$$

Basic Reaction

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$$

$$K_b = 1.8 \times 10^{-5} = \frac{[NH_4^+]\,[OH^-]}{[NH_3]}$$

Acidic Reaction

$$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons NH_{3(aq)} + H_3O^+{}_{(aq)}$$

$$K_a = 5.6 \times 10^{-10} = \frac{[NH_3]\,[H_3O^+]}{[NH_4^+]}$$

Relationship Between K_a and K_b

$K_w = K_a \times K_b$

$$\left(\frac{[NH_4^+][OH^-]}{[NH_3]}\right)\left(\frac{[NH_3][H_3O^+]}{[NH_4^+]}\right) = \quad [OH^-][H_3O^+] = K_w$$

$pK_w = pK_a + pK_b$ (@25°C)

Example: pK_a & pK_b

$pK_w = pK_a + pK_b$ (@25°C)

An acid was dissolved in water and the pK_b value associated with its conjugate base was found to be 15.5 at 25°C.

1. What is the pK_a value for the acid? $pK_a = -1.5$
2. Is the solution acidic or basic? Acidic
3. Is the conjugate base a stronger or a weaker base than water?

The conjugate base is a weaker base than water. Water and the conjugate base compete for protons and water wins every time (strong acid - HNO_3).

Example: Calculating pH for a Weak Acid Solution

Find the pH of a 2.5 M solution of HF.

$$HF_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + F^-_{(aq)} \qquad (\mathrm{p}K_a = 3.14)$$

Step 1: Find K_a

Step 2: Find $[H_3O^+]$ using an ICE chart

Step 3: Find pH using $[H_3O^+]$

	$[HF]$	$[H_3O^+]$	$[F^-]$
I			
C			
E			

Example: Calculating pH for a Weak Acid Solution

Find the pH and % ionization of a 1.5 M HCN solution.

$$\mathrm{HCN}_{(aq)} + \mathrm{H_2O}_{(l)} \rightleftharpoons \mathrm{H_3O^+}_{(aq)} + \mathrm{CN^-}_{(aq)} \quad (\mathrm{p}K_\mathrm{a} = 9.31)$$

Step 1: Find K_a	**Step 3: Find pH using $[H_3O^+]$**
Step 2: Find $[H_3O^+]$ using an ICE chart	**Step 4: Find % ionization.**

	$[HCN]$	$[H_3O^+]$	$[CN^-]$
I			
C			
E			

Example: K_a and pH with % Ionization

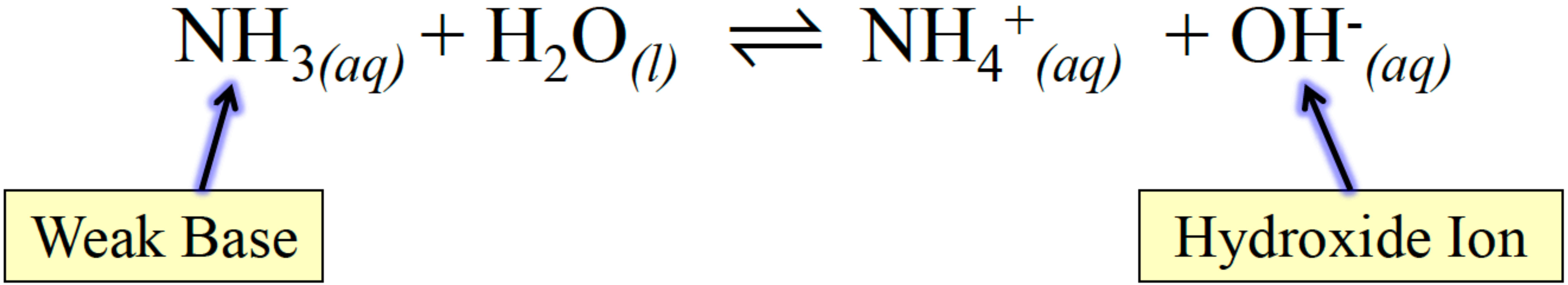
Acetic acid, $HC_2H_3O_2$, experiences 0.767% ionization in a 0.300 M solution at 25°C. Find K_a for acetic acid at 25°C and the pH of the solution.

Step 1: Find K_a	**Step 2: Find pH**

	$[HC_2H_3O_2]$ $\rightleftharpoons$	$[H^+]$ +	$[C_2H_3O_2^-]$
I			
C			
E			

Weak Bases

- Weak bases react with water to produce hydroxide ions, OH^-.



$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4{}^+{}_{(aq)} + OH^-{}_{(aq)}$$

Weak Base

Hydroxide Ion

The equilibrium for weak base reactions normally lies to the left.

Example: Calculating pH for Weak Base Solutions

- Find the pH of a 3.5 M solution of NH_3.

$pK_b = 4.74$

$$NH_{3(aq)} + H_2O_{(l)} \rightleftharpoons NH_4{}^+{}_{(aq)} + OH^-{}_{(aq)}$$

Step 1: Find K_b	**Step 3: Find pOH using $[OH^-]$**
Step 2: Find $[OH^-]$ using an ICE chart	**Step 4: Find pH**

Polyprotic Acids

- Polyprotic acids can donate more than one H^+ in a solution.

e.g. H_2SO_4 and H_2CO_3

Sulfuric acid

Carbonic Acid

$$H_2CO_{3(aq)} \rightleftharpoons H^+_{(aq)} + HCO_3^-{}_{(aq)} \quad (K_{a1} = 4.3 \text{ x } 10^{-7})$$

$$HCO_3^-{}_{(aq)} \rightleftharpoons H^+_{(aq)} + CO_3^{2-}{}_{(aq)} \quad (K_{a2} = 5.6 \text{ x } 10^{-11})$$

- Always use K_{a1} to calculate $[H^+]$ and pH.
  - Most of the H^+ ions come from the first ionization.
  - The H^+ ions from the first ionization drive the equilibrium(s) for the other ionizations to the left.

8.4 Acid-Base Reactions & Buffers

- Strong Acid-Strong Base
- Weak Acid-Strong Base
- Strong Acid-Weak Base
- Weak Acid-Weak Base

Acid-Base Reactions

- Reactions between acids and bases are called *neutralization reactions*.
- If a strong acid or base is involved, $K_{eq} > 1$, the reaction goes to completion and a one way arrow is used.
- In a weak acid-weak base reaction, $K_{eq} < 1$, a state of equilibrium is established and two way arrows are used.

Strong Acid - Strong Base Reactions

- Strong acids and strong bases experience 100% ionization.
- H^+ and OH^- ions combine to form H_2O.
  - These reactions go to completion.
  - $[H^+][OH^-] = K_w = 1 \times 10^{-14}$ at 25°C.
- The other parts of these acids and bases act as spectator ions.
- The pH can be determined from the excess reactant.

Example: Strong Acid - Strong Base Reactions

- 250.0 mL of 0.35 M HCl is mixed with 150.0 mL of 0.50 M NaOH. Write the net ionic equation for the reaction and find the pH of the final solution. Assume that the volumes are additive.

Step 1. Find the excess reactant.

$$0.2500\ \text{L} \times \frac{0.35\ \text{mol H}^+}{1.0\ \text{L}} = 0.088\ \text{mol H}^+$$

$$0.1500\ \text{L} \times \frac{0.50\ \text{mol OH}^-}{1.0\ \text{L}} = 0.075\ \text{mol OH}^-$$

Step 2. Find $[\text{H}^+]$

$$\frac{0.088\ \text{mol H}^+ - 0.075\ \text{mol OH}^-}{0.2500\ \text{L} + 0.1500\ \text{L}} = 0.033M\ \text{H}^+$$

Weak Acid - Strong Base Reactions

- Strong bases completely dissociate.
- Each OH^- ion rips an H^+ ion off of a weak acid molecule.
- This produces water and the conjugate base of the weak acid.

- Equal volumes of 1.0 M HNO_2 and 1.0 M KOH are mixed. Write the net ionic equation for the reaction and identify the aqueous species that have the highest concentrations at equilibrium.

Example: Weak Acid - Strong Base Reactions

- 55 mL of 2.5 M NaOH is added to 55 mL of 2.0 M HF. Assuming the volumes are additive, find the pH of the final solution at 25°C.

Step 1: Find the limiting reactant	**Step 3: Find pOH of the final solution**
Step 2: Find $[OH^-]$ in the final solution	**Step 4: Find pH of the final solution**

Henderson - Hasselbalch Equation

$$\mathrm{pH} = \mathrm{p}K_a + \log\frac{[\mathrm{A}^-]}{[\mathrm{HA}]}$$

or

$$\mathrm{pH} = \mathrm{p}K_a + \log\frac{[\mathrm{Base}]}{[\mathrm{Acid}]}$$

$\mathrm{p}K_\mathrm{a} = -\log_{10}K_\mathrm{a}$

$[\mathrm{A}^-]$ = molarity of conjugate base

$[\mathrm{HA}]$ = molarity of weak acid (initial molarity)

$$\log\frac{[\mathrm{Base}]}{[\mathrm{Acid}]}$$

- Must be a conjugate acid-base pair
- A weak base and its conjugate acid

K_a and pK_a - Remember!

$$\mathrm{p}K_a = -\log K_a$$

and

$$K_a = 10^{-\mathrm{p}K_a}$$

Example: Weak Acid - Strong Base

- 20.0 mL of 1.5 M KOH is added to 50.0 mL of 3.0 M CH_3COOH. $pK_a = 4.8$ at 25°C for CH_3COOH. Find the pH of the final solution.

Step 1: Find the limiting reactant

Step 3: Find pH of the final solution

Step 2: Find $[CH_3COO^-]$ and $[CH_3COOH]$ in the final solution

KOH is the limiting reactant. When the weak acid is the excess reactant, a buffered solution is created and the pH is calculated using the Henderson-Hasselbalch equation.

Strong Acid - Weak Base Reactions

- Weak Bases will accept protons from a strong acid.

- Weak Bases may be nitrogen containing compounds such as NH_3 or the conjugate base of weak acids.

- **Example**: Equal volumes of 2.0 M NH_3 and 2.0 M HI are mixed. Write the net ionic equation for the reaction and identify the aqueous species that have the highest concentrations at equilibrium.

$$NH_{3(aq)} + H_3O^+{}_{(aq)} \rightarrow NH_4^+{}_{(aq)} + H_2O_{(l)}$$

I^- and NH_4^+ will have the highest concentrations at equilibrium.

Example: Strong Acid - Weak Base Reactions

- 125 mL of 2.5 M HCl is added to 115 mL of 2.2 M NH_3. Find the pH of the final solution at 25°C.

Step 1: Find the limiting reactant

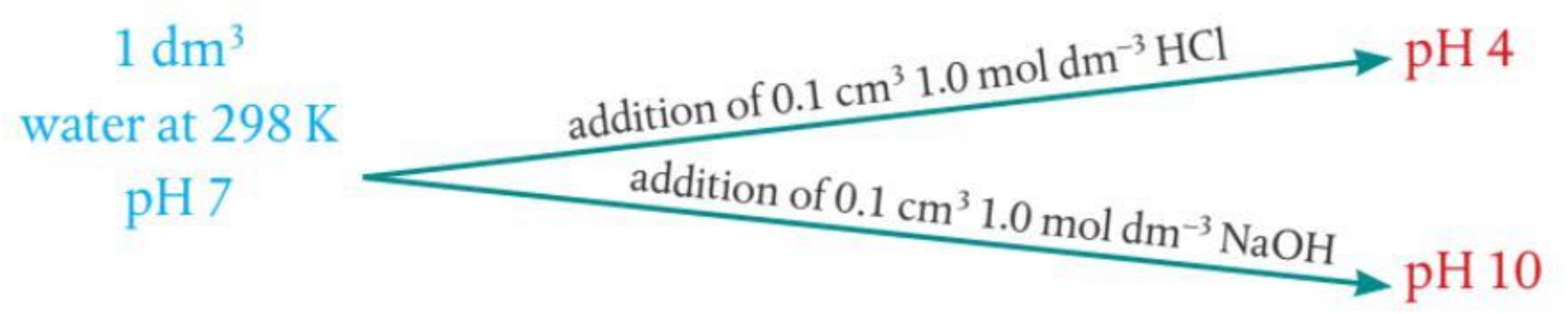
Step 2: Find $[H_3O^+]$ in the final solution.

Step 3: Find pH of the final solution

NH_3 is the limiting reactant. When the strong acid is the excess reactant, the pH can be found from the moles of excess H_3O^+ and the total volume of the final solution.

Buffer Solutions

- Something that acts to reduce the impact of one thing on another (a shock absorber of sorts)
- A *buffer solution* is resistant to changes in pH with the addition of **small amounts** of acid or base.



How Buffers Work

- 2 main types:
  - *acidic* - maintain the pH at a value less than 7
  - *basic* - maintain the pH at a value more than 7
- Mixture of 2 solutions
  - each contain a conjugate acid-base pair

Acidic Buffers

- Made by mixing a weak acid with a solution of its salt of a strong alkali
  - i.e. $CH_3COOH_{(aq)}$ with $NaCH_3COO$
    - weak acid salt of weak acid with strong alkali
  - Two equilibria exist in this buffer
    - $CH_3COOH \Leftrightarrow CH_3COO^- + H^+$
    - $NaCH_3COO \Leftrightarrow Na^+ + CH_3COO^-$

Has high concentrations of both acid and base - ready to react with either H^+ or OH^-

Response to Added Acids & Bases

- Adding acid -
  - $CH_3COO^- + H^+ \Leftrightarrow CH_3COOH$
- Adding base
  - $CH_3COOH + OH^- \Leftrightarrow CH_3COO^- + H_2O$

Basic Buffers

- Made by mixing a weak base with a solution of its salt of a strong acid
  - i.e. NH_3 with NH_4Cl
    - $NH_3 + H_2O \Leftrightarrow NH_4^+ + OH^-$
    - $NH_4Cl \rightarrow NH_4^+ + Cl^-$

Has high concentrations of both acid and base - ready to react with either H^+ or OH^-

Response to Added Acids & Bases

- Addition of acid
  - H^+ combines with the base NH_3 to form NH_4^+
- Addition of base
  - OH^- combines with the acid NH_4^+ to form NH_3 and water, removing most of OH^-

How to Make a Buffer Solution

- Start with a weak acid whose pK_a is very close to the desired pH of the buffer (or weak base whose pK_b is close to the desired pOH).
- Then, react this acid or base with enough strong alkali or strong acid so that one half of it is converted into salt.

	CH_3COOH	NaOH	CH_3COONa	H_2O
I	2	1	0	0
C	-1	-1	1	1
E	1	0	1	1

End up with equal amounts of weak acid and salt

Weak Acid - Weak Base Reactions

- Generally do not go to completion.
- Acids and bases are mostly undissociated.
- We write these as proton transfer reactions.
- Equal volume of 0.2 M aqueous ammonia and 0.2 M acetic acid are mixed.

$$CH_3COOH_{(aq)} + NH_{3(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + NH_4{}^+{}_{(aq)}$$

A two way arrow is used, as a state of equilibrium is established.

8.5 Acid-Base Titrations

- Endpoint
- Equivalence Point
- Half-Equivalence Point

Acid-Base Titrations

Acid-base titrations are often used to determine the concentration of an acid or a base in a solution.

- A strong acid of known concentration (titrant) is added to a base of unknown concentration (analyte), which is mixed with an indicator; or,

- *A strong base of known concentration (titrant) is added to an acid of unknown concentration (analyte), which is mixed with an indicator.*
    - The indicator changes color to signal the arrival of the ***Endpoint*** .

Acid-Base Titrations

Endpoint

- When the correct indicator is chosen, the endpoint is very close to the ***Equivalence Point***.

- ***Equivalence Point***
  - At the equivalence point, the number of moles of titrant added is equal to the number of moles of analyte that were originally present.

$$M_aV_a = M_bV_b$$

Three Types of Acid-Base Titrations

1. Strong acid - strong base
2. Weak acid - strong base
3. Weak base - strong acid

Strong Acid & Strong Base

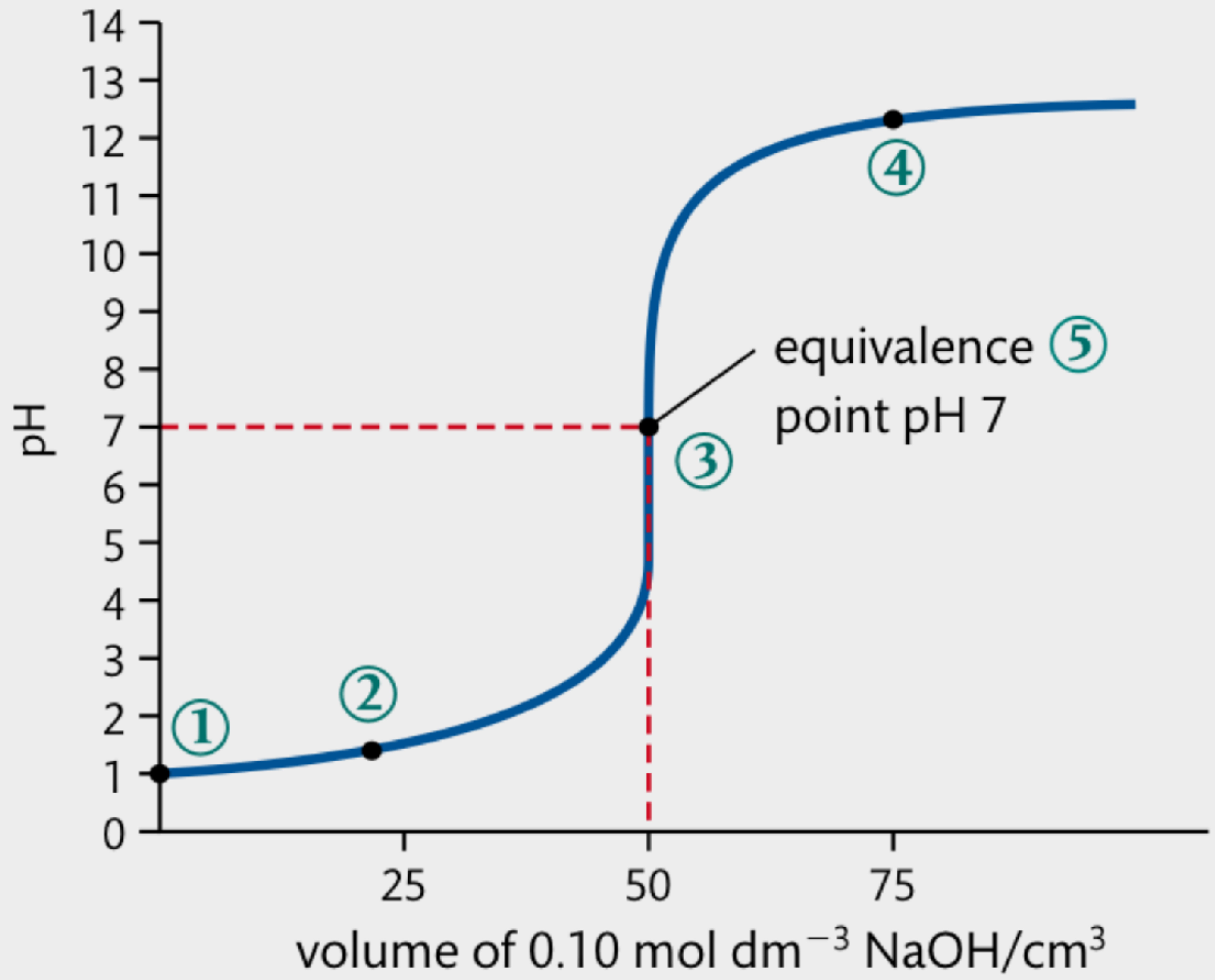
(1) Initial pH = 1

(2) pH changes only gradually until equivalence.

(3) Very sharp jump in pH at equivalence: from 3 to 11

(4) After equivalence, the curve flattens out at a high value.

(5) The pH at equivalence is 7



Weak Acid & Strong Base

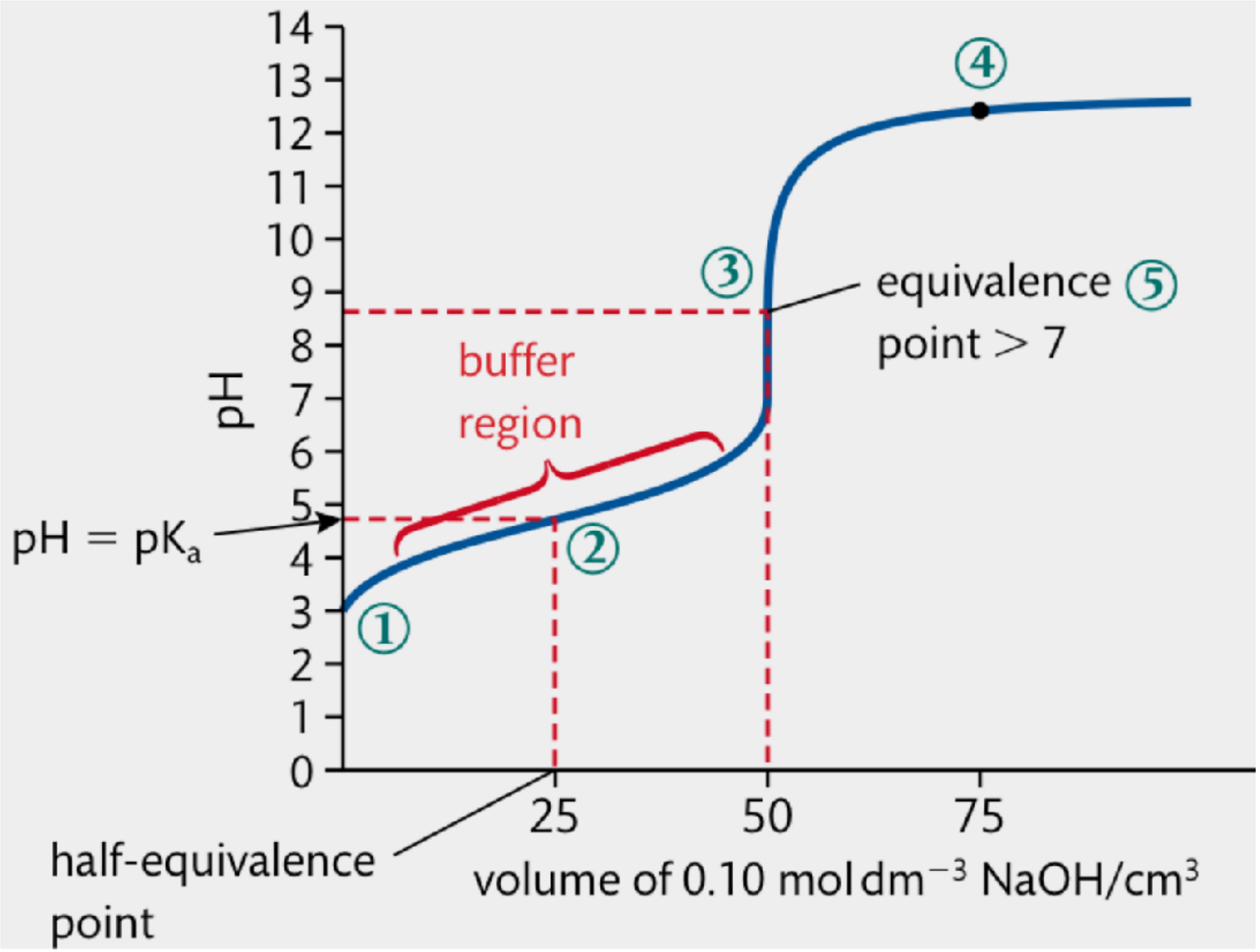
(1) Initial pH fairly high

(2) pH stays relatively constant until equivalence - labeled as **buffer region**

(3) Jump in pH at equivalence from about pH 7-11, which is not as much of a jump as for a SA-SB titration

(4) After equivalence, the curve flattens out at a high value (SB)

(5) pH at equivalence is > 7



Strong Acid & Weak Base

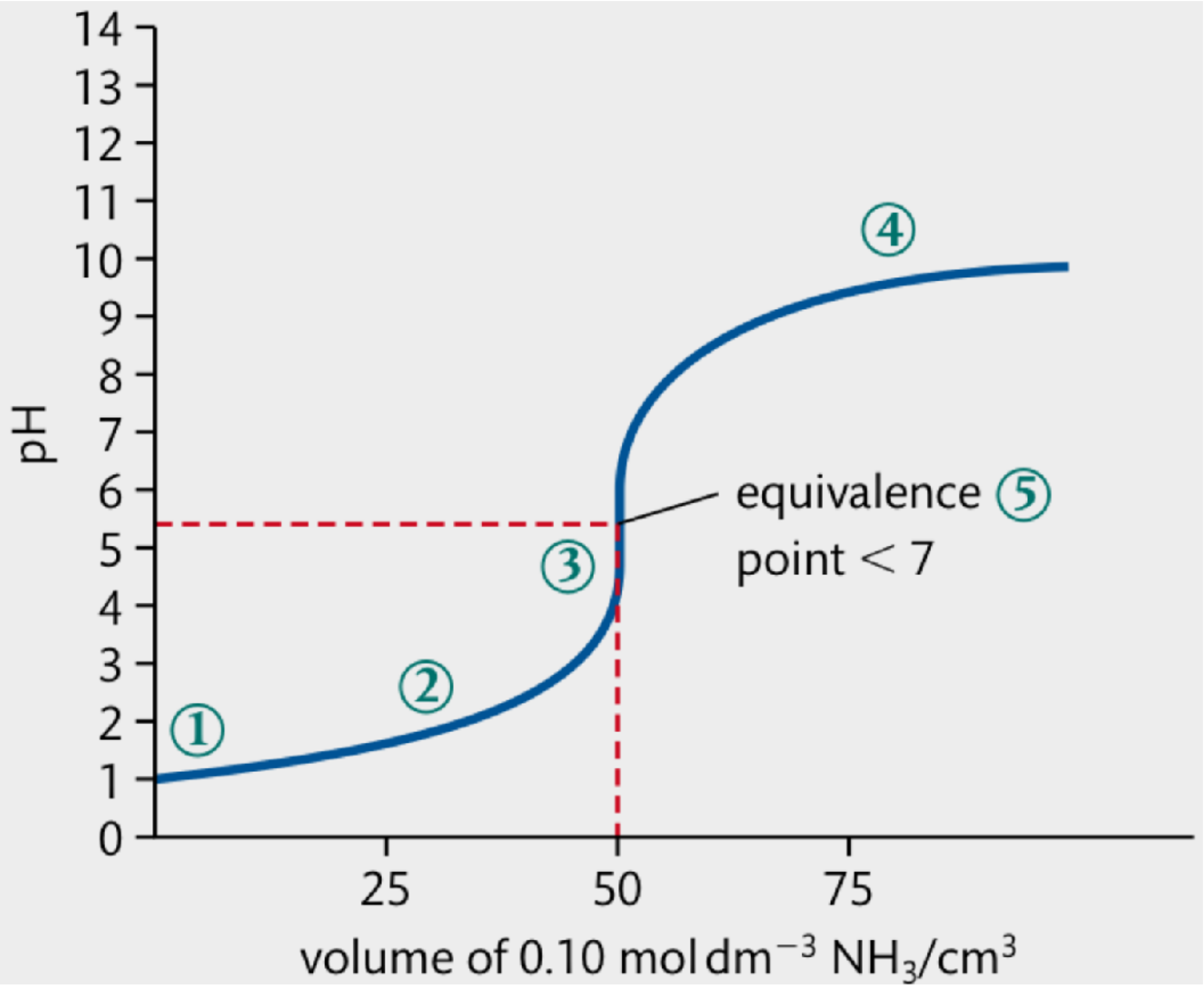
(1) Initial pH = 1 (SA)

(2) pH stays relatively constant through the buffer region to equivalence.

(3) Jump in pH at equivalence from about 3.0-7.0

(4) After equivalence, the curve flattens out at a fairly low pH (WB)

(5) pH at equivalence < 7



Example: Strong Acid - Strong Base Titration

30 mL of 0.50 M HCl was titrated with 0.50 M NaOH. Assume that volumes are additive and the final temperature of all solutions is 25°C.

a) Find the initial pH of 0.5 M HCl	**c) Find pH after 30 mL of NaOH was added.**
b) Find pH after 15 mL of NaOH was added.	**d) Find the pH after 45 mL of NaOH was added.**

The net ionic equation

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

Example: Weak Acid - Strong Base Titration

30 mL of 0.50 M CH_3COOH was titrated with 0.50 M NaOH.

a) Determine the volume of 0.50 M NaOH that must be delivered to reach the half equivalence point.

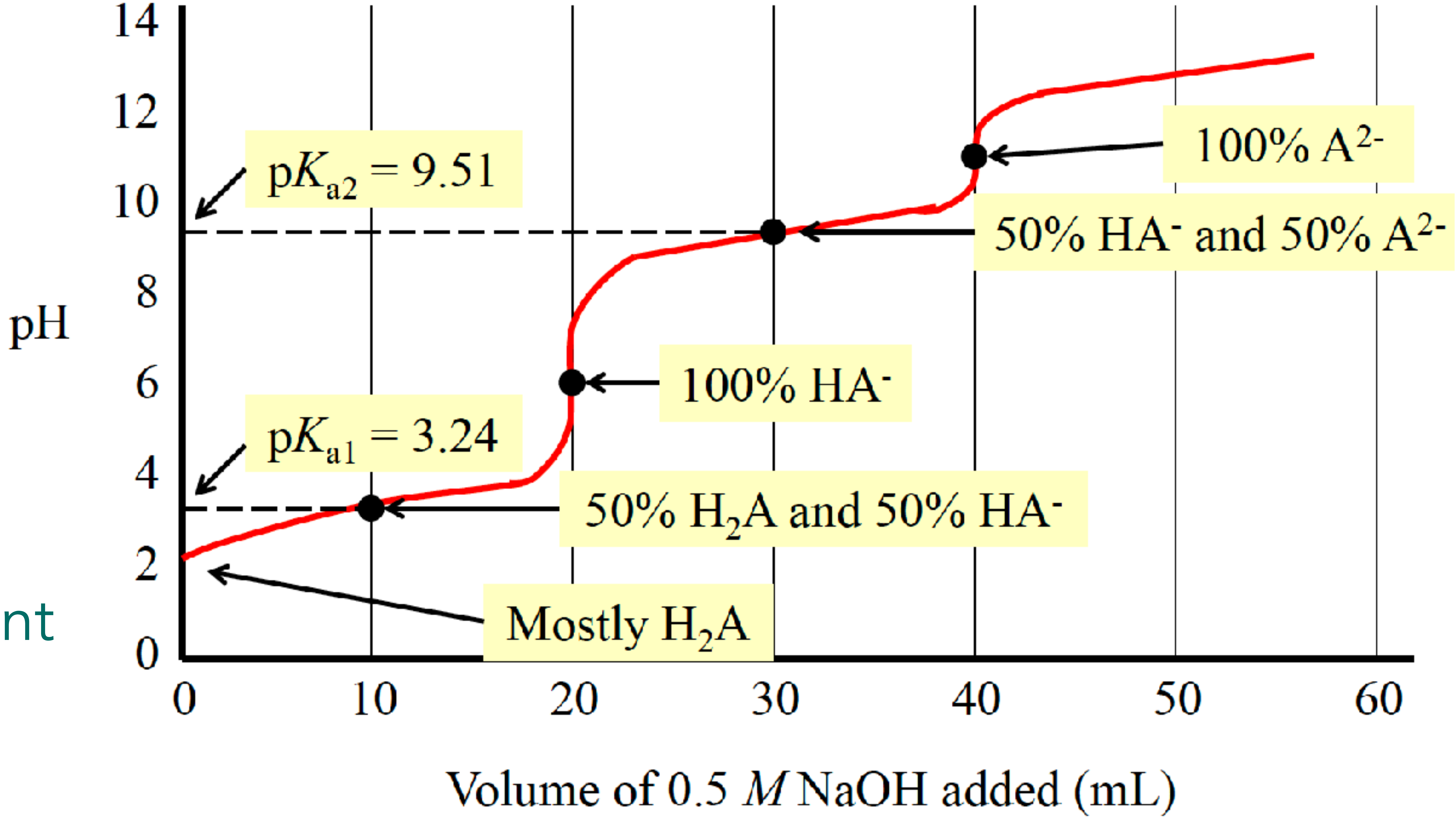
b) The pH at the half equivalence point is measured to be 4.74. Find pK_a for CH_3COOH.

c) Identify the main species that are present in the solution being titrated at the half equivalence point.

Titration Curves for Polyprotic Acids

The following can be determined:

- the number of acidic protons
- the pK$_a$ value for each acidic proton of a weak polyprotic acid
- The major species present at any point along the curve.



8.7 pH and pK_a
8.8 Properties of Buffers
8.9 Henderson-Hasselbalch Equation
8.10 Buffer Capacity

- Buffered Solutions
- Acid-Base Indicators

The Common Ion Effect is Used to Create Buffer Solutions

Buffers resist changes in pH

- When small quantities of strong acids or strong bases are added to a buffered solution, the changes in pH are small.

Buffer

- A weak acid and its salt (extra conjugate base)
- or… a weak base with its salt (extra conjugate acid)

Buffer Capacity

Buffering Capacity

- The ability of a buffered solution to accept protons or hydroxide ions without significant change in pH
- Higher concentrations of HA and A- in the solution result in greater buffering capacities.

$$\text{pH} = \text{p}K_a + \log\frac{[\text{A}^-]}{[\text{HA}]}$$

$$\text{When } \frac{[\text{A}^-]}{[\text{HA}]} = 1,\ \log\frac{[\text{A}^-]}{[\text{HA}]} = 0 \text{ and pH} = \text{p}K_a$$

This is helpful when choosing a conjugate acid-base pair. If a buffer needs to have a certain pH, one would choose a weak acid with a pK$_a$ value that is very close to the desired pH.

How Acid-Base Indicators Work

- In a titration, there is a large shift in pH as the equivalence point is passed.
- This large shift in pH causes the indicator's equilibrium to shift and its color to change.
- The color starts to change when $[HIn] \cong [In^-]$
- The color change experienced by different indicators occurs over different pH ranges.
  - You must select an indicator that changes color at a pH that is as close to possible to the pH at the equivalence point.

How Acid-Base Indicators Work

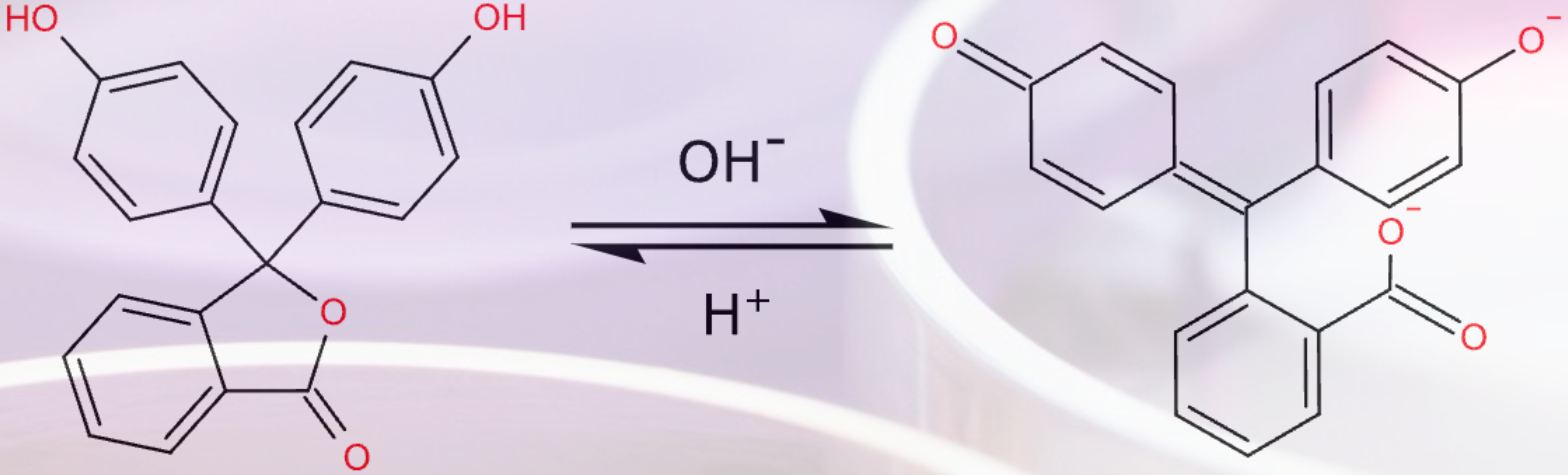
$$\mathrm{H}In_{(aq)} \rightleftharpoons \mathrm{H}^{+}_{(aq)} + In^{-}_{(aq)}$$

one color

another color

Colorless
pH<8.3

Pink colored
pH>10.5



Benzoid form of
Phenolphthalein
(acidic conditions)

Quinonoid form of
Phenolphthalein
(basic conditions)

How Acid-Base Indicators Work

$$\underset{\text{one color}}{\mathrm{H}In_{(aq)}} \rightleftharpoons \mathrm{H}^+_{(aq)} + \underset{\text{another color}}{In^-_{(aq)}}$$

When $[\mathrm{H}In] = [In^-]$…

$$\mathrm{pH} = \mathrm{p}K_\mathrm{a} + \log\frac{[In^-]}{[\mathrm{H}In]}$$

$$\mathrm{pH} = \mathrm{p}K_\mathrm{a}$$

Choose the indicator that has:

$\mathrm{p}K_{\mathrm{a(indicator)}} = \mathrm{pH}_{\mathrm{(equivalence\ point)}}$

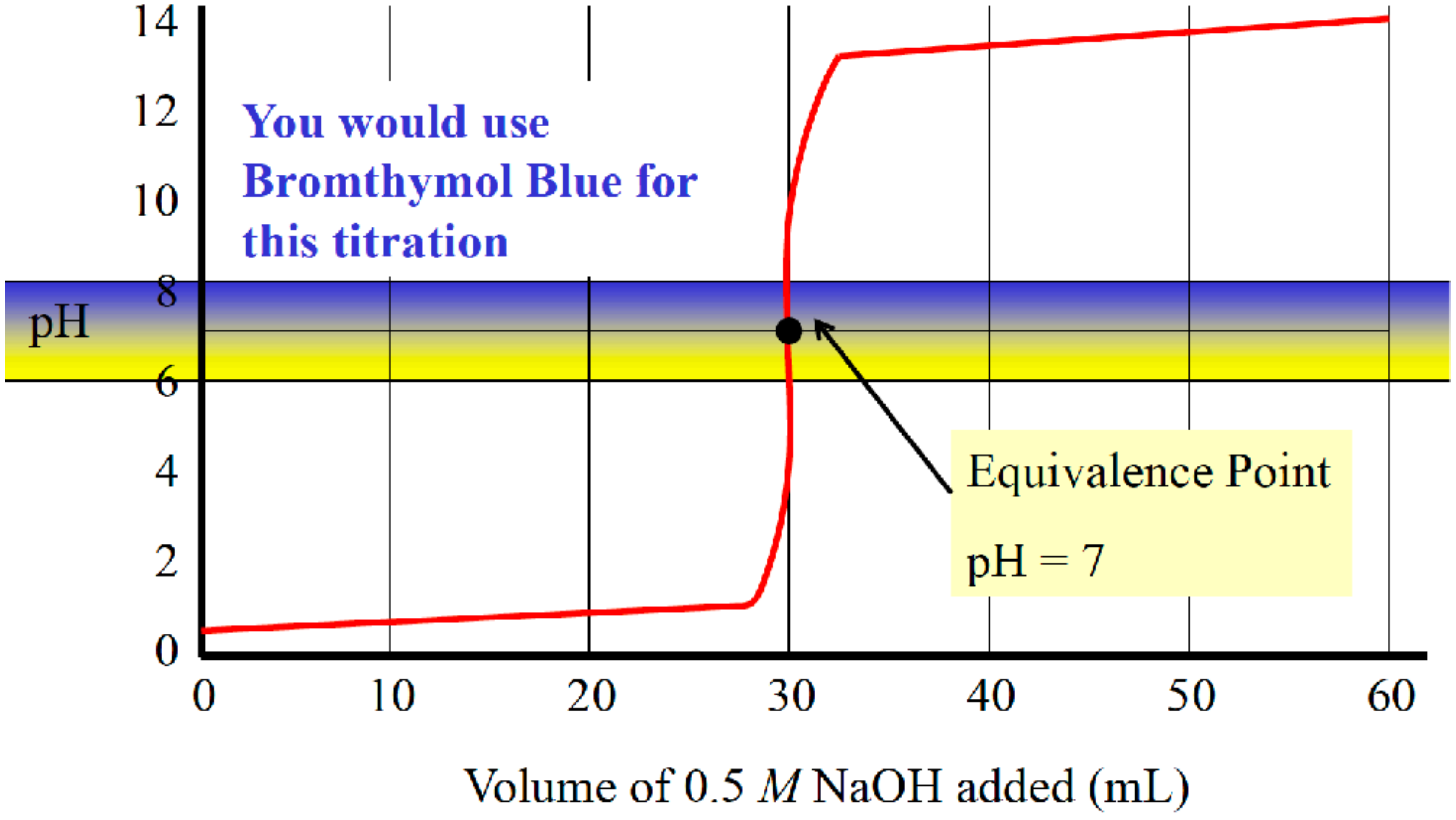
Choosing an Indicator

Indicator	pK_a
Methyl Orange	3.7
Methyl Red	5.1
Bromcresol Purple	6.3
Bromthymol Blue	7.0
Phenolphthalein	9.4
Alizarin Yellow	11.0

Example: Choosing an Indicator

Which indicator would you use when titrating HCl with NaOH? (Equivalence point occurs when pH = 7.0)

Indicator	pK_a
Methyl Orange	3.7
Methyl Red	5.1
Bromcresol Purple	6.3
Bromthymol Blue	7.0
Phenolphthalein	9.4
Alizarin Yellow	11.0



You would use Bromthymol Blue for this titration