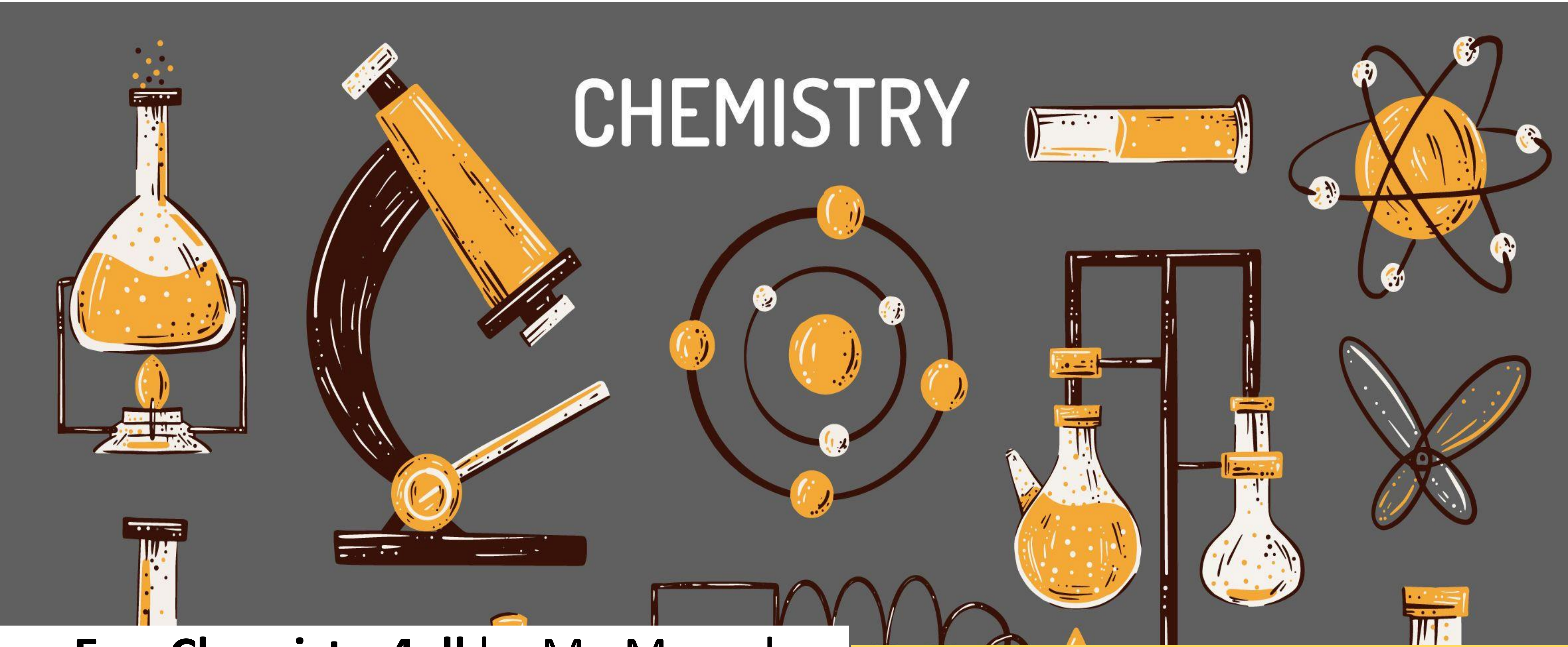


**EasyChemistry4all** by Mr. Mouad

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**00971557903129**

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# Inspire Chemistry
# Module 17
## "Acids & Bases"

0-14

## Lesson 3: "Hydrogen Ions and pH"

$H^+$

# Learning Outcomes:

- **Explain** pH and pOH.
- **Relate** pH and pOH to the ion product constant for water.
- **Calculate** the pH and pOH of aqueous solutions.

# Focus Question

What are pH and pOH?

**MAINIDEA** **pH and pOH are logarithmic scales that express the concentrations of hydrogen ions and hydroxide ions in aqueous solutions.**

$[H_3O^+]$

pH is related to $[H^+]$ in a solution.

pOH is related to $[OH^-]$ in a solution.

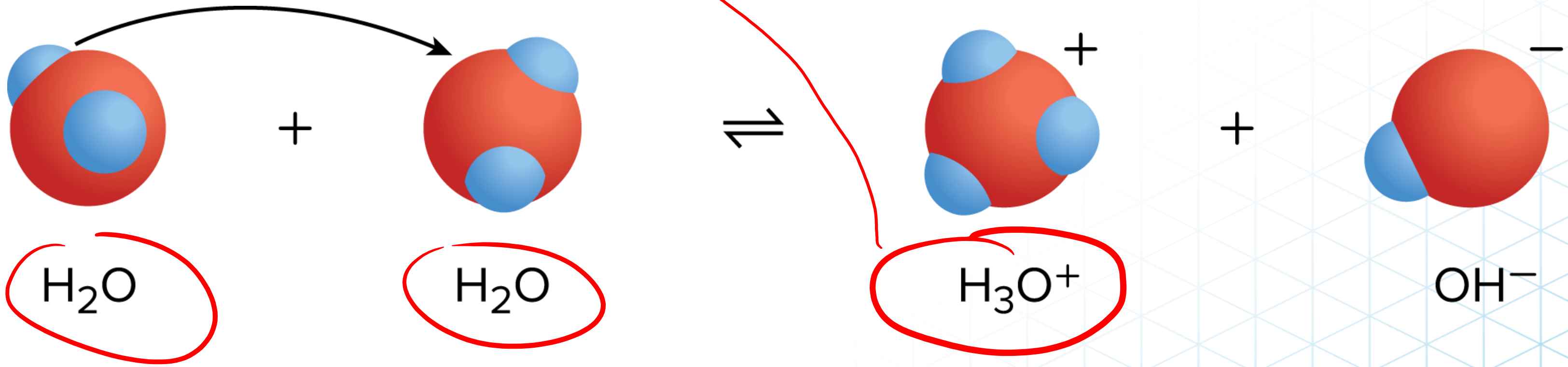
# Ion Product Constant for Water

lesson 1 $[H^+] = [OH^-]$

- Pure water contains **equal concentrations of $H^+$ and $OH^-$** ions produced by self-ionization.
- The equation for the equilibrium can be simplified as follows.

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

- In the self-ionization of water, one water molecule acts as an acid, and the other acts as a base.



# Ion Product Constant for Water

- The **ion product constant for water,** $K_w$ is the value of the equilibrium constant expression for the self-ionization of water.

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

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

# Ion Product Constant for Water

- **Fact from experiments:** With pure water at 298 K (25 °C), **both [H$^+$] and [OH$^-$] are equal to 1.0 × 10$^{-7}$*M*.**

$$K_w = [H^+][OH^-] = (1.0\times10^{-7})(1.0\times10^{-7})$$

$$K_w = 1.0\times10^{-14}$$

- According to Le Châtelier's Principle, as [H$^+$] goes up, [OH$^-$] must go down, and vice versa (والعكس صحيح). **This happens so that the value of K$_w$ will not change.**

# EXAMPLE 1

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**CALCULATE $[H^+]$ AND $[OH^-]$ USING $K_w$** At 298 K, the $H^+$ ion concentration in a cup of coffee is $1.0 \times 10^{-5} M$. What is the $OH^-$ ion concentration in the coffee? Is the coffee acidic, basic, or neutral?

**Known**

$[H^+] = 1.0 \times 10^{-5} M$

$K_w = 1.0 \times 10^{-14}$

**Unknown**

$[OH^-]$ = ? mol/L

$[H^+] > [OH^-]$ acidic

$[OH^-] > [H^+]$ basic solution

## 2 SOLVE FOR THE UNKNOWN

Use the ion product constant expression.

$K_w = [H^+][OH^-]$ — **State the ion product expression.**

$[OH^-] = \frac{K_w}{[H^+]}$ — **Solve for $[OH^-]$.**

$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}} = 1.0 \times 10^{-9}$ mol/L — **Substitute $K_w = 1.0 \times 10^{-14}$. Substitute $[H^+] = 1.0 \times 10^{-5} M$ and solve.**

Because $[H^+] > [OH^-]$, **the coffee is acidic.**

Remember!

**In acids: $[H^+]>[OH^-]$**

**In Bases: $[H^+]<[OH^-]$**

# CALCULATE [$H^+$] AND [$OH^-$] USING $K_W$

**IN-CLASS EXAMPLE**

*Use with Example Problem 1.*

## Problem

At 298 K, the $H^+$ ion concentration in a cup of coffee is $1.0 \times 10^{-5} M$. What is the $OH^-$ ion concentration in the coffee? Is the coffee acidic, basic, or neutral?

## Response

*ANALYZE THE PROBLEM*

You are given the concentration of the $H^+$ ion, and you know that $K_w$ equals $1.0 \times 10^{-14}$. You can use the ion product constant expression to solve for [$OH^-$]. Because [$H^+$] is greater than $1.0 \times 10^{-7}$, you can predict that [$OH^-$] will be less than $1.0 \times 10^{-7}$.

**KNOWN**

[$H^+$] = $1.0 \times 10^{-5} M$

$K_w = 1.0 \times 10^{-14}$

**UNKNOWN**

[$OH^-$] = ? mol/L

*SOLVE FOR THE UNKNOWN*

Use the ion product constant expression.

- State the ion product expression.

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

- Solve for [$OH^-$].

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

- Substitute $K_w = 1.0 \times 10^{-14}$. Substitute [$H^+$] = $1.0 \times 10^{-5} M$ and solve.

$$[OH^-] = \frac{1.0 \times 10{-14}}{1.0 \times 10{-5}} = 1.0 \times 10^{-9} \text{ mol/L}$$

Because [$H^+$] >[$OH^-$], **the coffee is acidic**.

*EVALUATE THE ANSWER*

The answer is correctly stated with two significant figures because [$H^+$] and $K_w$ each have two significant figures. As predicted, [$OH^-$] is less than $1.0 \times 10^{-7}$ mol/L.

22. The concentration of either the $H^+$ ion or the $OH^-$ ion is given for four aqueous solutions at 298 K. For each solution, calculate $[H^+]$ or $[OH^-]$. State whether the solution is acidic, basic, or neutral.

a. $[H^+] = 1.0 \times 10^{-13} M$

b. $[OH^-] = 1.0 \times 10^{-7} M$

c. $[OH^-] = 1.0 \times 10^{-3} M$

d. $[H^+] = 4.0 \times 10^{-5} M$

23. **Challenge** Calculate the number of $H^+$ ions and the number of $OH^-$ ions in 300 mL of pure water at 298 K.

22 The concentration of either the $H^+$ ion or the $OH^-$ ion is given for four aqueous solutions at 298 K. For each solution, calculate $[H^+]$ or $[OH^-]$. State whether the solution is acidic, basic, or neutral.

**a.** $[H^+] = 1.0 \times 10^{-13} M$

$$K_w = [H^+][OH^-]$$
$$1.0\times10^{-14} = (1.0\times10^{-13})[OH^-]$$
$$\frac{1.0\times10^{-14}}{1.0\times10^{-13}} = \frac{(1.0\times10^{-13})[OH^-]}{1.0\times10^{-13}}$$
$$[OH^-] = 1.0\times10^{-1} M$$

$[OH^-] > [H^+]$, the solution is basic.

**b.** $[OH^-] = 1.0 \times 10^{-7} M$

$$K_w = [H^+][OH^-]$$
$$\frac{1.0\times10^{-14}}{1.0\times10^{-7}} = \frac{[H^+](1.0\times10^{-7})}{1.0\times10^{-7}}$$
$$[H^+] = 1.0\times10^{-7} M$$

$[OH^-] = [H^+]$, the solution is netural.

**c.** $[OH^-] = 1.0 \times 10^{-3} M$

$$K_w = [H^+][OH^-]$$
$$1.0\times10^{-14} = [H^+](1.0\times10^{-3})$$
$$\frac{1.0\times10^{-14}}{1.0\times10^{-3}} = \frac{[H^+](1.0\times10^{-3})}{1.0\times10^{-3}}$$
$$[H^+] = 1.0\times10^{-11} M$$

$[OH^-] > [H^+]$, the solution is basic.

**d.** $[H^+] = 4.0 \times 10^{-5} M$

$$K_w = [H^+][OH^-]$$
$$1.0\times10^{-14} = (4.0\times10^{-5})[OH^-]$$
$$\frac{1.0\times10^{-14}}{1.0\times10^{-5}} = \frac{(4.0\times10^{-5})[OH^-]}{4.0\times10^{-5}}$$
$$[OH^-] = 2.5\times10^{-10} M$$

$[H^+] > [OH^-]$, the solution is acidic.

23 **Challenge** Calculate the number of $H^+$ ions and the number of $OH^-$ ions in 300 mL of pure water at 298 K.

At 298 K, $[H^+] = [OH^-] = 1.0\times10^{-7} M$

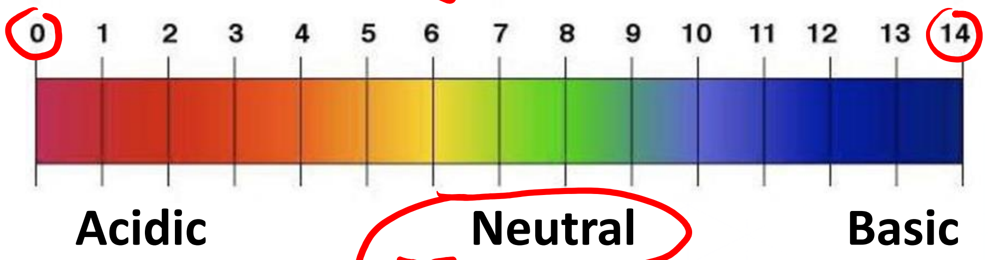
$$\text{Mol } H^+ = \frac{1.0\times10^{-7}\text{ mol}}{1\text{L}} \times \frac{1\text{ L}}{1000\text{ mL}} \times 300\text{ mL}$$
$$= 3.0\times10^{-8}\text{ mol}$$
$$3.0\times10^{-8}\text{ mol } H^+\text{ ions} \times \frac{6.02\times10^{-3}\ H^+\text{ ions}}{1\text{ mol } H^+}$$
$$= 1.8\times10^{16}\ H^+\text{ ions}$$

Number of $H^+$ = number of $OH^-$

$$= 1.8\times10^{16}\text{ ions}$$

# pH and pOH

## pH Scale




=7

0 - 7 acidic

7 - 14 basic

## Rules of pH:

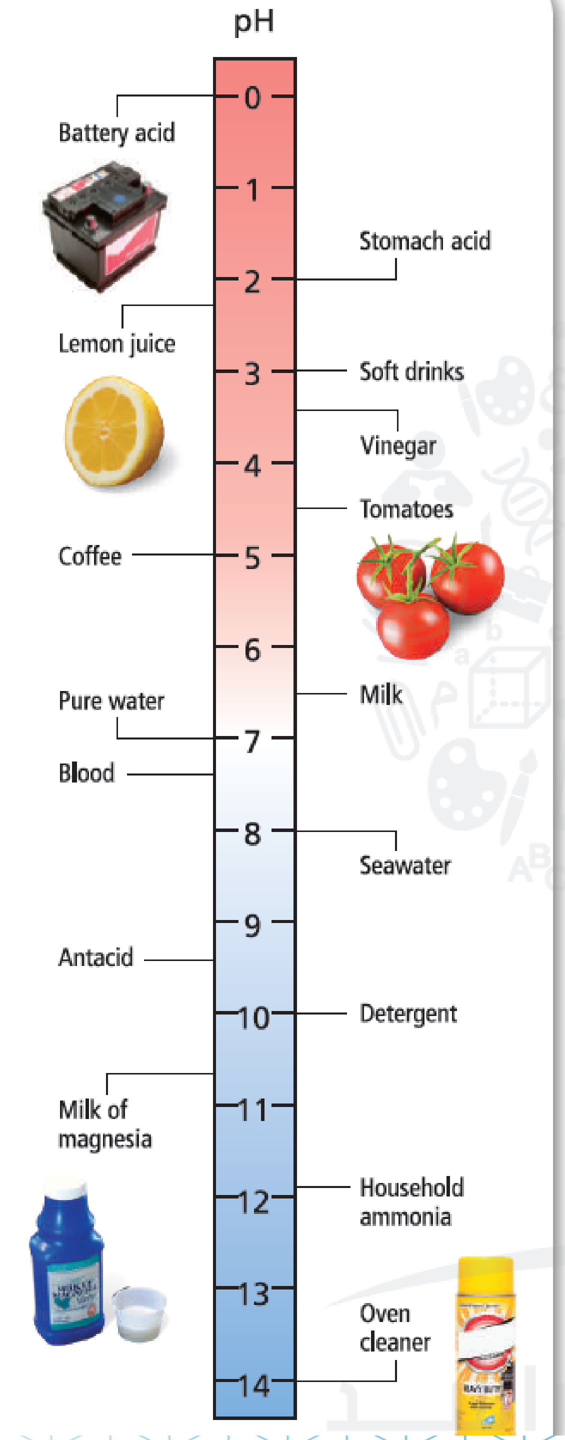
pH = 1

When the pH value is closer to zero ⟶ It is a strong acid

When the pH value is 7 ⟶ It is a neutral compound

When the pH value is closer to 14 ⟶ It is a strong base

pH = 12.5



# pH and pOH

$[H^+] = 3 \times 10^{-9}$ M

0.0000…

- Concentrations of $H^+$ and $OH^-$ ions are often small numbers expressed in scientific notation.
- pH and pOH are easier ways to express these small concentrations.
- **pH** is the **negative logarithm** of **the hydrogen ion concentration** of a solution.

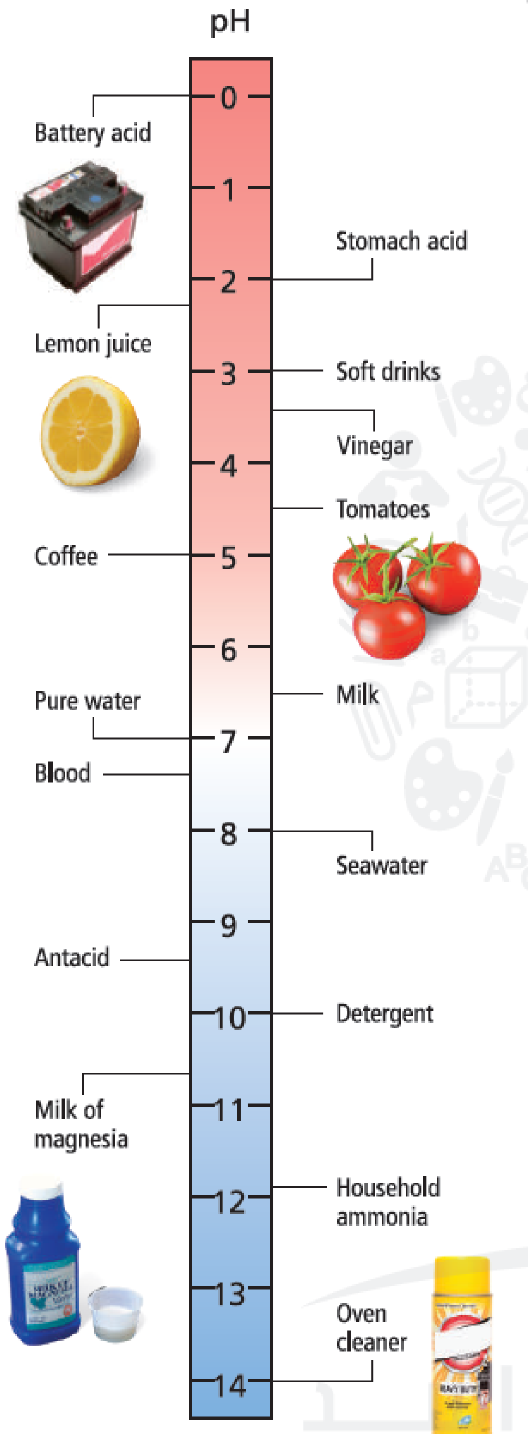
$[H^+] = 1 \times 10^{-5} \Rightarrow pH = -\log(1 \times 10^{-5}) = 5$

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

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

$pH = 5$

$[H^+] = 10^{-pH} = 10^{-5} = 1 \times 10^{-5}$ M



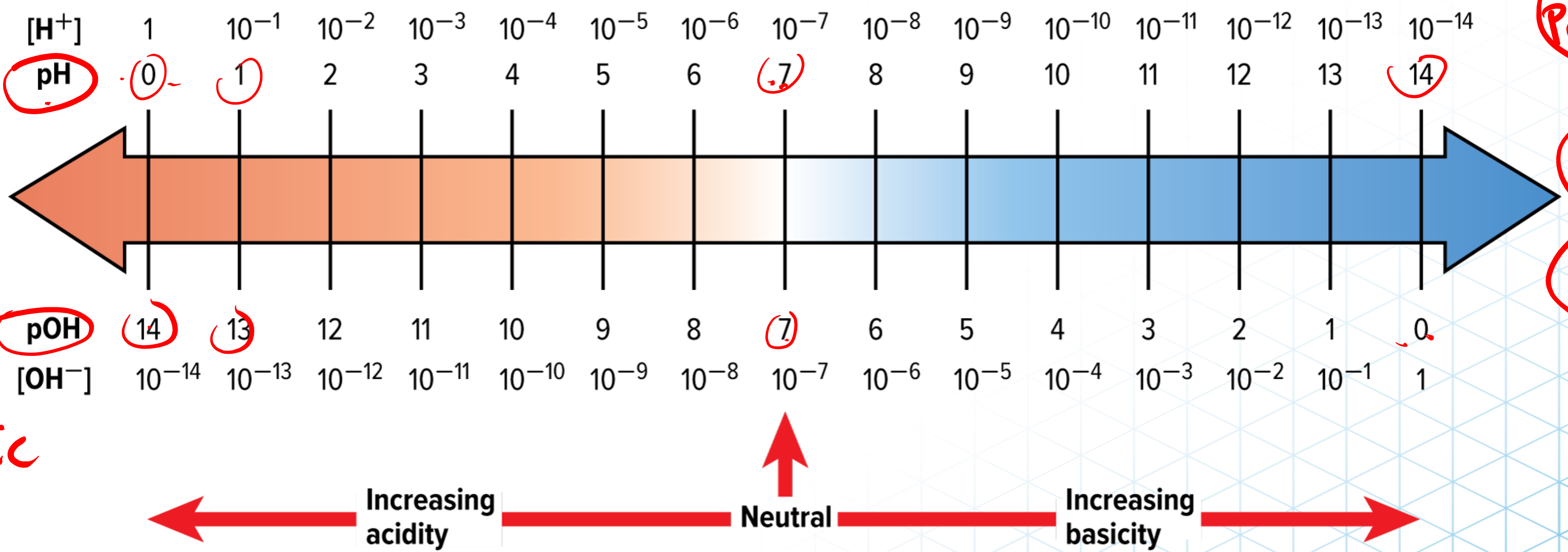
# pH and pOH

- The **pOH** of a solution is the **negative logarithm of the hydroxide ion** concentration.

$$pOH = -\log [OH^-]$$

- **The sum of pH and pOH is 14.**

**$pH + pOH = 14$**



# Problem types:

CALCULATE PH FROM $[H^+]$

CALCULATE pOH AND pH FROM $[OH^-]$

CALCULATE $[H^+]$ AND $[OH^-]$ FROM pH

**What you need!!**

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

$pOH = -\log[OH^-]$

$[OH^-] \times [H^+] = 10^{-14}$

$pH + pOH = 14$

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

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

## Molarity & pH of strong and weak acids

CALCULATE $K_a$ FROM pH

## EXAMPLE 2

**CALCULATE PH FROM [$H^+$]** What is the pH of a neutral solution at 298 K?

pH = 7

Neutral solutions : $[H^+] = [OH^-] = 1 \times 10^{-7}$

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

$pH = -\log(1 \times 10^{-7}) = 7$

# EXAMPLE 2

**CALCULATE PH FROM $[H^+]$** What is the pH of a neutral solution at 298 K?

## 1 ANALYZE THE PROBLEM

In a neutral solution at 298 K, $[H^+] = 1.0 \times 10^{-7} M$. You must find the negative log of $[H^+]$.

**Known**

$[H^+] = 1.0 \times 10^{-7}\ M$

**Unknown**

$pH = ?$

## 2 SOLVE FOR THE UNKNOWN

$pH = -\log [H^+]$ — **State the equation for pH.**

$pH = -\log (1.0 \times 10^{-7})$ — **Substitute $[H^+] = 1.0 \times 10^{-7} M$.**

The pH of the neutral solution at 298 K is 7.00.

## APPLICATIONS



**24.** Calculate the pH of solutions having the following ion concentrations at 298 K.

$= -\log[H^+]$

**a.** $[H^+] = 1.0 \times 10^{-2} M$ **b.** $[H^+] = 3.0 \times 10^{-6} M$

**25.** Calculate the pH of aqueous solutions with the following $[H^+]$ at 298 K.

**a.** $[H^+] = 0.0055 M$ **b.** $[H^+] = 0.000084 M$

**26.** **Challenge** Calculate the pH of a solution having $[OH^-] = 8.2 \times 10^{-6} M$.

(1)

$pOH = -\log[OH^-] = -\log(8.2 \times 10^{-6}) = 5.08$

$pH + pOH = 14$

$pH = 14 - pOH = 14 - 5.08$

$pH = 8.9$

(2)

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

$[H^+] = \frac{1 \times 10^{-14}}{[OH^-]} = \frac{1 \times 10^{-14}}{8.2 \times 10^{-6}}$

$= 1.22 \times 10^{-9} M$

$pH = -\log[H^+] = -\log(1.22 \times 10^{-9})$

$= 8.9$

**24.** Calculate the pH of solutions having the following ion concentrations at 298 K.

**a.** $[H^+] \times 1.0 \times 10^{-2}M$

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

$pH = -\log(1.0 \times 10^{-2})$

$pH = 2.00$

**b.** $[H^+] = 3.0 \times 10^{-6}M$

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

$pH = -\log(3.0 \times 10^{-6})$

$pH = 5.52$

**25.** Calculate the pH of aqueous solutions having the following $[H^+]$ at 298 K.

**a.** $[H^+] = 0.0055M$

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

$pH = -\log 0.0055$

$pH = 2.26$

**b.** $[H^+] = 0.000084M$

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

$pH = -\log 0.000084$

$pH = 4.08$

**26. Challenge** Calculate the pH of a solution having $[OH^-] = 8.2 \times 10^{-6}M$.

$[OH^-] = 8.2 \times 10^{-6}M$

$K_w = [H^+][OH^-] \times [H^+](8.2 \times 10^{-6})$

$$[H^+] = \frac{1.0 \times 10^{-14}}{8.2 \times 10^{-6}} = 1.2 \times 10^{-9}$$

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

$pH = -\log(1.2 \times 10^{-9})$

$pH = 8.92$

## EXAMPLE 3

**CALCULATE pOH AND pH FROM $[OH^-]$** In **Figure 16,** a cow is being fed straw and hay that has been treated with ammonia. The addition of ammonia to animal feed promotes protein growth in the animal. Another use of ammonia is as a household cleaner, which is an aqueous solution of ammonia gas. A typical cleaner has a hydroxide-ion concentration of $4.0 \times 10^{-3} M$. Calculate the pOH and pH of a cleaner at 298 K.

**Known**

$[OH^-] = 4.0 \times 10^{-3} M$

**Unknown**

pOH = ?

pH = ?

## EXAMPLE 3

**CALCULATE pOH AND pH FROM $[OH^-]$** In **Figure 16,** a cow is being fed straw and hay that has been treated with ammonia. The addition of ammonia to animal feed promotes protein growth in the animal. Another use of ammonia is as a household cleaner, which is an aqueous solution of ammonia gas. A typical cleaner has a hydroxide-ion concentration of $4.0 \times 10^{-3}M$. Calculate the pOH and pH of a cleaner at 298 K.

| **Known** | **Unknown** |
|---|---|
| $[OH^-] = 4.0 \times 10^{-3}M$ | pOH = ? |
| | pH = ? |

### 2 SOLVE FOR THE UNKNOWN

$\mathrm{pOH} = -\log [OH^-]$ — **State the equation for pOH.**

$\mathrm{pOH} = -\log (4.0 \times 10^{-3})$ — **Substitute $[OH^-] = 4.0 \times 10^{-3}M$.**

The pOH of the solution is 2.40.

Use the relationship between pH and pOH to find the pH.

$\mathrm{pH} + \mathrm{pOH} = 14.00$ — **State the equation that relates pH and pOH.**

$\mathrm{pH} = 14.00 - \mathrm{pOH}$ — **Solve for pH.**

$\mathrm{pH} = 14.00 - 2.40 = 11.60$ — **Substitute pOH = 2.40.**

The pH of the solution is 11.60.

# EXAMPLE 4

**CALCULATE $[H^+]$ AND $[OH^-]$ FROM pH** What are $[H^+]$ and $[OH^-]$ in a healthy person's blood that has a pH of 7.40? Assume that the temperature of the blood is 298 K.

**Known**

pH = 7.40

**Unknown**

$[H^+]$ = ? mol/L

$[OH^-]$ = ? mol/L

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

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

3.98 ~ 4

① $[H^+] = 10^{-pH} = 10^{-7.4}$

$[H^+] = 3.98 \times 10^{-8}$ M ~ 4

$[OH^-] = \frac{1 \times 10^{-14}}{3.98 \times 10^{-14}} = 2.51 \times 10^{-7}$ M

② $pOH = 14 - pH = 14 - 7.4 = 6.6$

$[H^+] = 10^{-pH} = 10^{-7.4} = 3.98 \times 10^{-8}$ M

$[OH^-] = 10^{-pOH} = 10^{-6.6} = 2.51 \times 10^{-7}$ M

## EXAMPLE 4

**CALCULATE $[H^+]$ AND $[OH^-]$ FROM pH** What are $[H^+]$ and $[OH^-]$ in a healthy person's blood that has a pH of 7.40? Assume that the temperature of the blood is 298 K.

| **Known** | **Unknown** |
| --- | --- |
| pH = 7.40 | $[H^+]$ = ? mol/L |
| | $[OH^-]$ = ? mol/L |

### 2 SOLVE FOR THE UNKNOWN

Determine $[H^+]$.

$pH = -\log [H^+]$ **State the equation for pH.**

$-pH = \log [H^+]$ **Multiply both sides of the equation by −1.**

$[H^+] = \text{antilog}(-pH)$ **Take the antilog of each side to solve for $[H^+]$.**

$[H^+] = \text{antilog}(-7.40)$ **Substitute pH = 7.40.**

$[H^+] = 4.0 \times 10^{-8} M$ **A calculator shows that the antilog of −7.40 is $4.0 \times 10^{-8}$.**

5.94

## EXAMPLE 4

**CALCULATE $[H^+]$ AND $[OH^-]$ FROM pH** What are $[H^+]$ and $[OH^-]$ in a healthy person's blood that has a pH of 7.40? Assume that the temperature of the blood is 298 K.

$[H^+] = 4.0 \times 10^{-8} M$ — A calculator shows that the antilog of −7.40 is $4.0 \times 10^{-8}$.

The concentration of $H^+$ ions in the blood is $4.0 \times 10^{-8} M$.

Determine $[OH^-]$.

| | |
|---|---|
| $pH + pOH = 14.00$ | State the equation that relates pH and pOH. |
| $pOH = 14.00 - pH$ | Solve for pOH. |
| $pOH = 14.00 - 7.40 = 6.60$ | Substitute pH = 7.40. |
| $pOH = -\log [OH^-]$ | State the equation for pOH. |
| $-pOH = \log [OH^-]$ | Multiply both sides of the equation by −1. |
| $[OH^-] = \text{antilog}(-6.60)$ | Take the antilog of each side and substitute pOH = 6.60. |
| $[OH^-] = 2.5 \times 10^{-7} M$. | A calculator shows that the antilog of −6.60 is $2.5 \times 10^{-7}$. |

The concentration of $OH^-$ ions in the blood is $2.5 \times 10^{-7} M$.

## APPLICATIONS

30. Calculate $[H^+]$ and $[OH^-]$ in each of the following solutions.

a. Milk, pH = 6.50

b. Lemon juice, pH = 2.37

c. Milk of magnesia, pH = 10.50

d. Household ammonia, pH = 11.90

31. **Challenge** Calculate the $[H^+]$ and $[OH^-]$ in a sample of seawater with a pOH = 5.60.

$$pH = 14 - 5.6 = 8.4$$

$$[H^+] = 10^{-8.4} = 3.98 \times 10^{-9}\ M$$

$$[OH^-] = 10^{-5.6} = 2.51 \times 10^{-6}\ M$$

## APPLICATIONS

30. Calculate $[H^+]$ and $[OH^-]$ in each of the following solutions.
   - a. Milk, pH = 6.50
   - b. Lemon juice, pH = 2.37
   - c. Milk of magnesia, pH = 10.50
   - d. Household ammonia, pH = 11.90

31. **Challenge** Calculate the $[H^+]$ and $[OH^-]$ in a sample of seawater with a pOH = 5.60.

**30. Calculate** $[H^+]$ and $[OH^-]$ in each of the following solutions.

**a.** Milk, pH = 6.50

$[H^+]$ = antilog (−pH)

$[H^+]$ = antilog (−6.50) = $3.2 \times 10^{-7}$ *M*

pOH = 14.00 − pH = 14.00 − 6.50 = 7.5

$[OH^-]$ = antilog (−pOH)

$[OH^-]$ = (−7.50) = $3.2 \times 10^{-8}$ *M*

**b.** Lemon juice, pH = 2.37

$[H^+]$ = antilog (−pH)

$[H^+]$ = antilog (−2.37) = $4.3 \times 10^{-3}$ *M*

pOH = 14.00 − pH = 14.00 − 2.37 = 11.63

$[OH^-]$ = antilog (− pOH)

$[OH^-]$ = antilog (−11.63) = $2.3 \times 10^{-12}$ *M*

**c.** Milk of magnesia, pH = 10.50

$[H^+]$ = antilog (−pH)

$[H^+]$ = antilog (−10.50) = $3.2 \times 10^{-11}$ M

pOH = 14.00 − pH = 14.00 − 10.50 = 3.50

$[OH^-]$ = antilog (−3.50) = $3.2 \times 10^{-4}$ *M*

**d.** Household ammonia, pH 11.90

$[H^+]$ = antilog (−pH)

$[H^+]$ = antilog (−11.90) = $1.3 \times 10^{-12}$ *M*

pOH = 14.00 − pH = 14.00 −11.90 = 2.10

$[OH^-]$ = antilog (−2.10) = $7.9 \times 10{-}^{3}$ *M*

**31. Challenge** Calculate the $[H^+]$ and $[OH^-]$ in a sample of seawater with a pOH = 5.60.

$[OH^-]$ = antilog (−pOH)

$[OH^-]$ = antilog (−5.60) = $2.5 \times 10^{-6}$ *M*

pH = 14.00 − 5.60 = 8.40

$[H^+]$ = antilog (−8.40) = $4.0 \times 10^{-9}$ *M*

# pH and pOH

- For all strong monoprotic acids, the concentration of the acid is the concentration of $H^+$ ions.
- For all strong bases, the concentration of the base is the concentration of available $OH^-$ ions.
- Weak acids and weak bases only partially ionize, so $K_a$ and $K_b$ values must be used to calculate pH and pOH.
- Litmus paper or a pH meter with electrodes can be used to determine the pH of a solution.

# Molarity and the pH of strong acids

$$HCl(aq) \longrightarrow H^{+}(aq) + Cl^{-}(aq)$$

Every HCl molecule produces one $H^{+}$ ion. The bottle labeled *0.1M HCl* contains 0.1 mol of $H^{+}$ ions per liter and 0.1 mol of $Cl^{-}$ ions per liter. For all strong monoprotic acids, the concentration of the acid is the concentration of $H^{+}$ ions. Thus, you can use the molarity of the acid to calculate pH.

Strong Acid

$[HCl] = 1\times10^{-3} = [H^{+}]$

pH = ??

$pH = -\log[H^{+}] = -\log(1\times10^{-3})$

$= 3$

# Molarity and the pH of strong acids

strong $H_2A \rightarrow 2H^+ + A^-$

$[H_2A] = 1\times10^{-3} \Rightarrow [H^+] = 2 \times 1\times10^{-3} = 2\times10^{-3}\,M$

$pH = ??\quad -\log(2\times10^{-3}) = 2.69$

# Molarity and the pH of strong bases

0.1*M* solution of the strong base NaOH in **Figure 18.17** is fully ionized.

$$NaOH(aq) \longrightarrow Na^{+}(aq) + OH^{-}(aq)$$

One formula unit of NaOH produces one $OH^-$ ion. Thus, the concentration of the $OH^-$ ions is the same as the molarity of the solution, 0.1*M*.

pOH

$$Ca(OH)_2 \longrightarrow Ca^{2+} + 2OH^-$$

$$[Ca(OH)_2] = 0.1\ M$$

$$[OH^-] = 2 \times 0.1 = 0.2\ M$$

$$pOH = -\log(0.2)$$

# Molarity and the pH of strong bases

# Molarity & pH of weak acids

**CALCULATE $K_a$ FROM pH**

## EXAMPLE Problem 18.5

**Calculate $K_a$ from pH** Formic acid is used to process latex tapped from rubber trees into natural rubber. The pH of a 0.100*M* solution of formic acid (HCOOH) is 2.38. What is $K_a$ for HCOOH?

$$HCOOH_{(aq)} \rightleftharpoons 1\,H^+_{(aq)} + 1\,COOH^-_{(aq)}$$

$[H^+] = [COOH^-] = 4.168 \times 10^{-3}$

known

pH = 2.38

[HCOOH] = 0.1 M

unknown

Ka ??

$$K_a = \frac{[H^+][COOH^-]}{[HCOOH]}$$

$[H^+] = 10^{-pH} = 10^{-2.38} = 4.168 \times 10^{-3}$

$$Ka = \frac{(4.168 \times 10^{-3})(4.168 \times 10^{-3})}{0.1} = 1.73 \times 10^{-4}$$

# Molarity & pH of weak acids

$Q_{sp}$ $K_{sp}$

**CALCULATE $K_a$ FROM pH**

## EXAMPLE Problem 18.5

$1\,HCOOH \rightleftharpoons 1\,H^+ + 1\,COOH^-$

**Calculate $K_a$ from pH** Formic acid is used to process latex tapped from rubber trees into natural rubber. The pH of a 0.100*M* solution of formic acid (HCOOH) is 2.38. What is $K_a$ for HCOOH?

$[HCOOH]_i = 0.1\,M$

$[HCOOH]_{used} = 4.168 \times 10^{-3}$

$[H^+]_{eq} = 4.168 \times 10^{-3}$

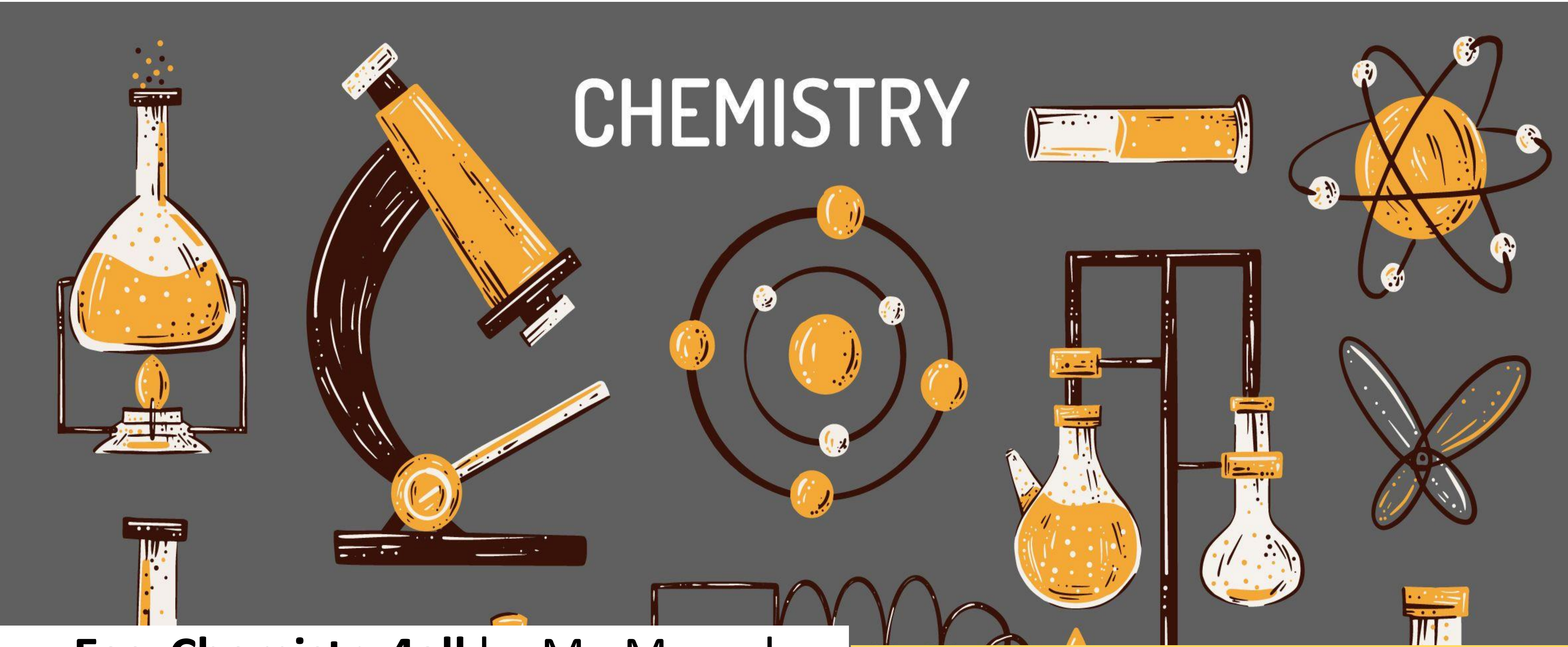
$[COOH^-]_{eq} = 4.168 \times 10^{-3}$

$[HCOOH]_{f,eq} = [HCOOH]_i - [HCOOH]_{used}$

$= 0.1 - 4.168 \times 10^{-3} = 0.096\,M$

$$K_a = \frac{[H^+][COOH^-]}{[HCOOH]_{f,eq}} = \frac{(4.168 \times 10^{-3})(4.168 \times 10^{-3})}{0.096} = 1.8 \times 10^{-4}$$

$K_a$

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