

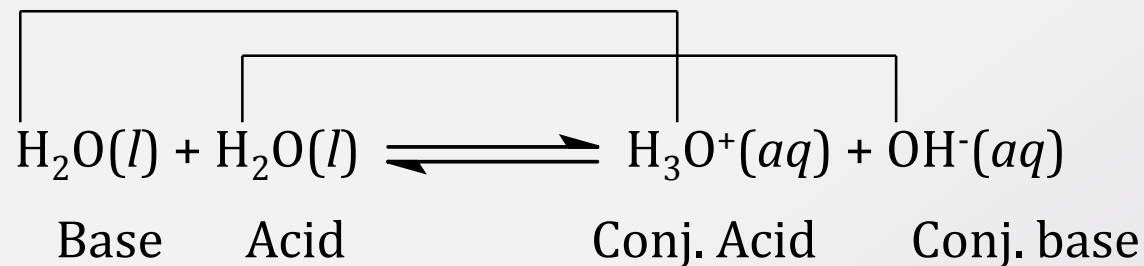
ACIDS AND BASES - 2

IONIZATION OF WATER, pH

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IONIZATION OF WATER

Self-Ionization of Water



We call the equilibrium constant the ion-product constant, K_w .

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] \text{ At } 25^\circ\text{C}, K_w = 1.0 \times 10^{-14}$$

As temperature increases, the value of K_w increases.

Solutions can be characterized as

$$\text{Acidic: } [\text{H}_3\text{O}^+] > 1.0 \times 10^{-7} \text{ M}$$

$$\text{Neutral: } [\text{H}_3\text{O}^+] = 1.0 \times 10^{-7} \text{ M}$$

$$\text{Basic: } [\text{H}_3\text{O}^+] < 1.0 \times 10^{-7} \text{ M}$$

EXAMPLE: CALCULATION OF CONCENTRATIONS

Calculate the hydronium and hydroxide ion concentration at 25°C in

a. 0.10 M HCl

b. 1.4×10^{-4} M Mg(OH)₂

Solution:

a. When HCl ionizes, it gives H⁺ and Cl⁻.

So $[H^+] = [Cl^-] = [HCl] = 0.10$ M.

b. When Mg(OH)₂ ionizes, it gives Mg²⁺ and 2 OH⁻.

So $2[OH^-] = 2(1.4 \times 10^{-4} \text{ M}) = 2.8 \times 10^{-4}$ M.

CALCULATING pH

$pH = -\log[H_3O^+]$

And inverse: $[H_3O^+] = 10^{-pH}$

- When pH = 7, the solution is neutral ($[H_3O^+] = 1.0 \times 10^{-7} M$).
- When pH < 7.00, the solution is acidic ($[H_3O^+] > 1.0 \times 10^{-7} M$).
- When pH > 7.00, the solution is basic ($[H_3O^+] < 1.0 \times 10^{-7} M$).

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| $[H_3O^+](M)$ | $-\log [H_3O^+]$ | pH | |
|-----------------------|-------------------------------|-------|---------|
| 0.10 | $-\log (1.0 \times 10^{-1})$ | 1.00 | ↑ |
| 0.010 | $-\log (1.0 \times 10^{-2})$ | 2.00 | |
| 1.0×10^{-3} | $-\log (1.0 \times 10^{-3})$ | 3.00 | |
| 1.0×10^{-4} | $-\log (1.0 \times 10^{-4})$ | 4.00 | |
| 1.0×10^{-5} | $-\log (1.0 \times 10^{-5})$ | 5.00 | |
| 1.0×10^{-6} | $-\log (1.0 \times 10^{-6})$ | 6.00 | |
| 1.0×10^{-7} | $-\log (1.0 \times 10^{-7})$ | 7.00 | |
| | | | Neutral |
| 1.0×10^{-8} | $-\log (1.0 \times 10^{-8})$ | 8.00 | ↓ |
| 1.0×10^{-9} | $-\log (1.0 \times 10^{-9})$ | 9.00 | |
| 1.0×10^{-10} | $-\log (1.0 \times 10^{-10})$ | 10.00 | |
| 1.0×10^{-11} | $-\log (1.0 \times 10^{-11})$ | 11.00 | |
| 1.0×10^{-12} | $-\log (1.0 \times 10^{-12})$ | 12.00 | |
| 1.0×10^{-13} | $-\log (1.0 \times 10^{-13})$ | 13.00 | |
| 1.0×10^{-14} | $-\log (1.0 \times 10^{-14})$ | 14.00 | |

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| Fluid | pH | Fluid | pH |
|--------------------------|---------|-------------------|-----------|
| Stomach acid | 1.0 | Saliva | 6.4–6.9 |
| Lemon juice | 2.0 | Milk | 6.5 |
| Vinegar | 3.0 | Pure water | 7.0 |
| Grapefruit juice | 3.2 | Blood | 7.35–7.45 |
| Orange juice | 3.5 | Tears | 7.4 |
| Urine | 4.8–7.5 | Milk of magnesia | 10.6 |
| Rainwater (in clean air) | 5.5 | Household ammonia | 11.5 |

EXAMPLE: CALCULATING pH

1) What is the pH of a solution that has a hydronium ion concentration of $6.5 \times 10^{-5}M$?

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log[6.54 \times 10^{-5}]$$

$$\text{pH} = 4.19$$

2) What is the hydronium ion concentration of a solution with pH 3.65?

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

$$[\text{H}_3\text{O}^+] = 10^{-3.65}$$

$$[\text{H}_3\text{O}^+] = 2.2 \times 10^{-4}$$

pH ,pOH AND K_w

- pOH can be calculated just like pH.

$$\text{pOH} = -\log[\text{OH}^-]$$

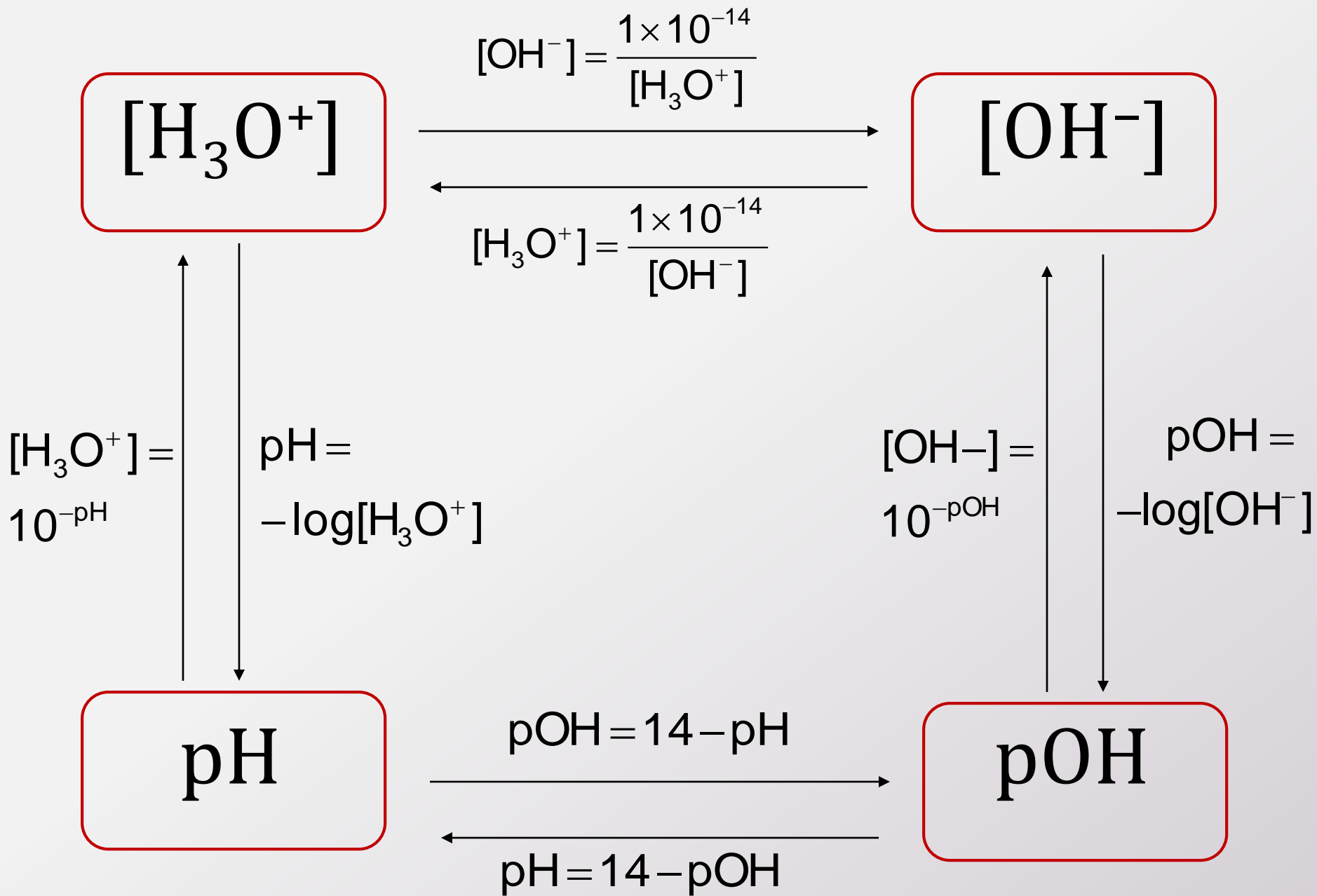
$$[\text{OH}^-] = 10^{-\text{pOH}}$$

- K_w is the hydronium ion and hydroxide ion concentrations; so $\text{p}K_w$ is the addition of pH and pOH.

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$\text{p}K_w = \text{pH} + \text{pOH} = 14.00$$

- Next slide has all relationships...



EXAMPLE - pOH

1) What is the pOH of a solution that has a hydroxide ion concentration of $4.3 \times 10^{-2}M$?

$$\text{pOH} = -\log[\text{OH}^-]$$

$$\text{pOH} = -\log[4.3 \times 10^{-2}]$$

$$\text{pOH} = 1.37$$

2) What is the hydroxide ion concentration of a solution with pOH 8.35?

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

$$[\text{OH}^-] = 10^{-8.35}$$

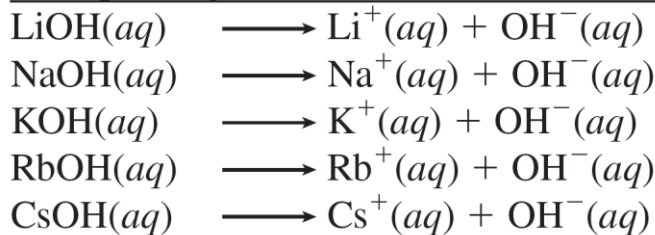
$$[\text{OH}^-] = 4.5 \times 10^{-9}$$

STRONG ACIDS AND BASES

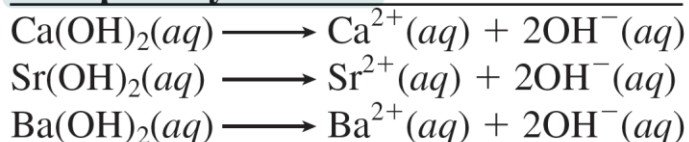
- All strong acids and bases ionize completely; so their concentrations can be used to calculate the pH and pOH.
- For diprotic acids, e.g. sulfuric acid only the first ionization reactions is measured; on the contrary dihydroxide bases will ionize to give two hydroxides in solution.

| Strong Acid | Ionization Reaction |
|--------------------|---|
| Hydrochloric acid | $\text{HCl}(aq) + \text{H}_2\text{O}(l) \longrightarrow \text{H}_3\text{O}^+(aq) + \text{Cl}^-(aq)$ |
| Hydrobromic acid | $\text{HBr}(aq) + \text{H}_2\text{O}(l) \longrightarrow \text{H}_3\text{O}^+(aq) + \text{Br}^-(aq)$ |
| Hydroiodic acid | $\text{HI}(aq) + \text{H}_2\text{O}(l) \longrightarrow \text{H}_3\text{O}^+(aq) + \text{I}^-(aq)$ |
| Nitric acid | $\text{HNO}_3(aq) + \text{H}_2\text{O}(l) \longrightarrow \text{H}_3\text{O}^+(aq) + \text{NO}_3^-(aq)$ |
| Chloric acid | $\text{HClO}_3(aq) + \text{H}_2\text{O}(l) \longrightarrow \text{H}_3\text{O}^+(aq) + \text{ClO}_3^-(aq)$ |
| Perchloric acid | $\text{HClO}_4(aq) + \text{H}_2\text{O}(l) \longrightarrow \text{H}_3\text{O}^+(aq) + \text{ClO}_4^-(aq)$ |
| Sulfuric acid | $\text{H}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l) \longrightarrow \text{H}_3\text{O}^+(aq) + \text{HSO}_4^-(aq)$ |

Group 1A hydroxides



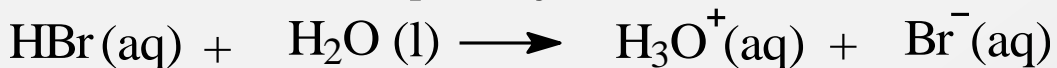
Group 2A hydroxides



EXAMPLE

- What is the pH of a 0.057 M solution of HBr?

Solution: HBr ionizes completely so the conc. of the acid can be used as H_3O^+ .



$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log[0.057]$$

$$\text{pH} = 1.24 \quad \text{Note: very low pH - strong acid.}$$

- What is the pOH of a solution of 0.034 M solution of Ca(OH)_2 ?

Solution: Ca(OH)_2 ionizes to give 2 mols of OH^- , so conc. has to be doubled.



$$[\text{OH}^-] = 0.034 \text{ M Ca(OH)}_2 \times \frac{2 \text{ mol OH}^-}{1 \text{ mol Ca(OH)}_2}$$

$$[\text{OH}^-] = 0.068 \text{ M}$$

$$\text{pOH} = -\log[\text{OH}^-]$$

$$\text{pOH} = -\log[0.068]$$

$$\text{pOH} = 1.17$$

KEY CONCEPTS

- Calculation of pH and pOH
- Calculation of hydronium and hydroxide ion conc.