

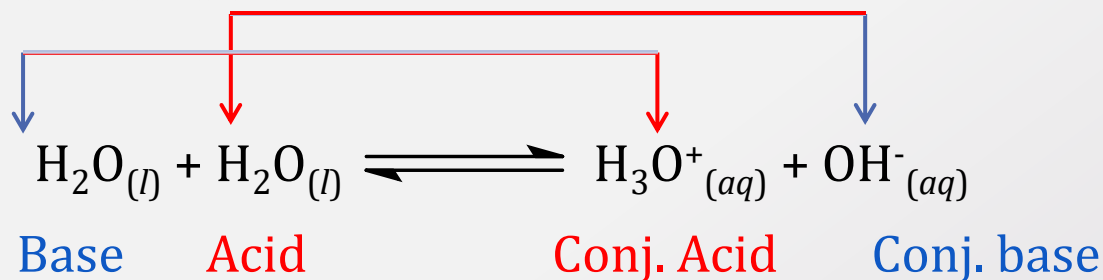
ACIDS AND BASES

2: IONIZATION OF WATER, pH CALCULATION

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IONIZATION OF WATER

If we look at self-ionization of water, where water is the acid and the base, then we get the reaction below.



We can write the equilibrium expression as: $K_c = [\text{H}_3\text{O}^+][\text{OH}^-]$
(Remember that liquids are not part of K_c expression).

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

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

At 25°C , $K_w = 1.0 \times 10^{-14}$ as shown by using the concentration of both hydronium and hydroxide ions as $1 \times 10^{-7} \text{ M}$.

$$K_w = [1 \times 10^{-7}][1 \times 10^{-7}]$$

IMPLICATION OF K_w

K_w is calculated as follows:

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

$$1.0 \times 10^{-14} = [1 \times 10^{-7}][1 \times 10^{-7}]$$

Solutions can be characterized as

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

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

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

EXAMPLE 1: CALCULATION OF CONCENTRATIONS

Calculate the concentrations at 25°C in

- a) Hydronium ions in 0.10 M HCl.
- b) Hydroxide ions in 1.4×10^{-4} M $\text{Mg}(\text{OH})_2$.
- c) Hydroxide ions in 2.30×10^{-7} M nitric acid.

Answer:

- a) When HCl ionizes, it gives H^+ and Cl^- .

$$\text{So } [\text{H}^+] = [\text{Cl}^-] = [\text{HCl}] = 0.10 \text{ M}$$

- b) When $\text{Mg}(\text{OH})_2$ ionizes, it gives Mg^{2+} and 2 OH^- .

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

- c) Nitric acid dissociates completely. So $[\text{H}_3\text{O}^+]$ is 2.30×10^{-7} M. Use the K_w equation to find $[\text{OH}^-]$.

$$1.0 \times 10^{-14} = [\text{H}_3\text{O}^+][\text{OH}^-]$$

$$[\text{OH}^-] = 1.0 \times 10^{-14} / 2.30 \times 10^{-7} = 4.3 \times 10^{-8} \text{ M}$$

CALCULATING pH

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

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

- When $\text{pH} = 7$, the solution is neutral ($[\text{H}_3\text{O}^+] = 1.0 \times 10^{-7} \text{ M}$).
- When $\text{pH} < 7.00$, the solution is acidic ($[\text{H}_3\text{O}^+] > 1.0 \times 10^{-7} \text{ M}$).
- When $\text{pH} > 7.00$, the solution is basic ($[\text{H}_3\text{O}^+] < 1.0 \times 10^{-7} \text{ M}$).

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	7.40
Orange juice	3.5	Milk of magnesia	10.6
Rain water	5.5	Household ammonia	11.5

$[\text{H}_3\text{O}^+] \text{ (M)}$	$-\log [\text{H}_3\text{O}^+]$	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
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

EXAMPLE 2: 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?

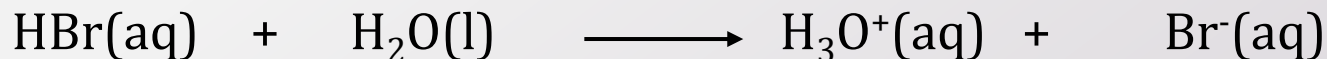
$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

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

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

- 3) What is the pH of a 0.0570 M solution of HBr?

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



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

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

$$\text{pH} = 1.24$$

pH ,pOH AND K_w

- pOH can be calculated just like pH.

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

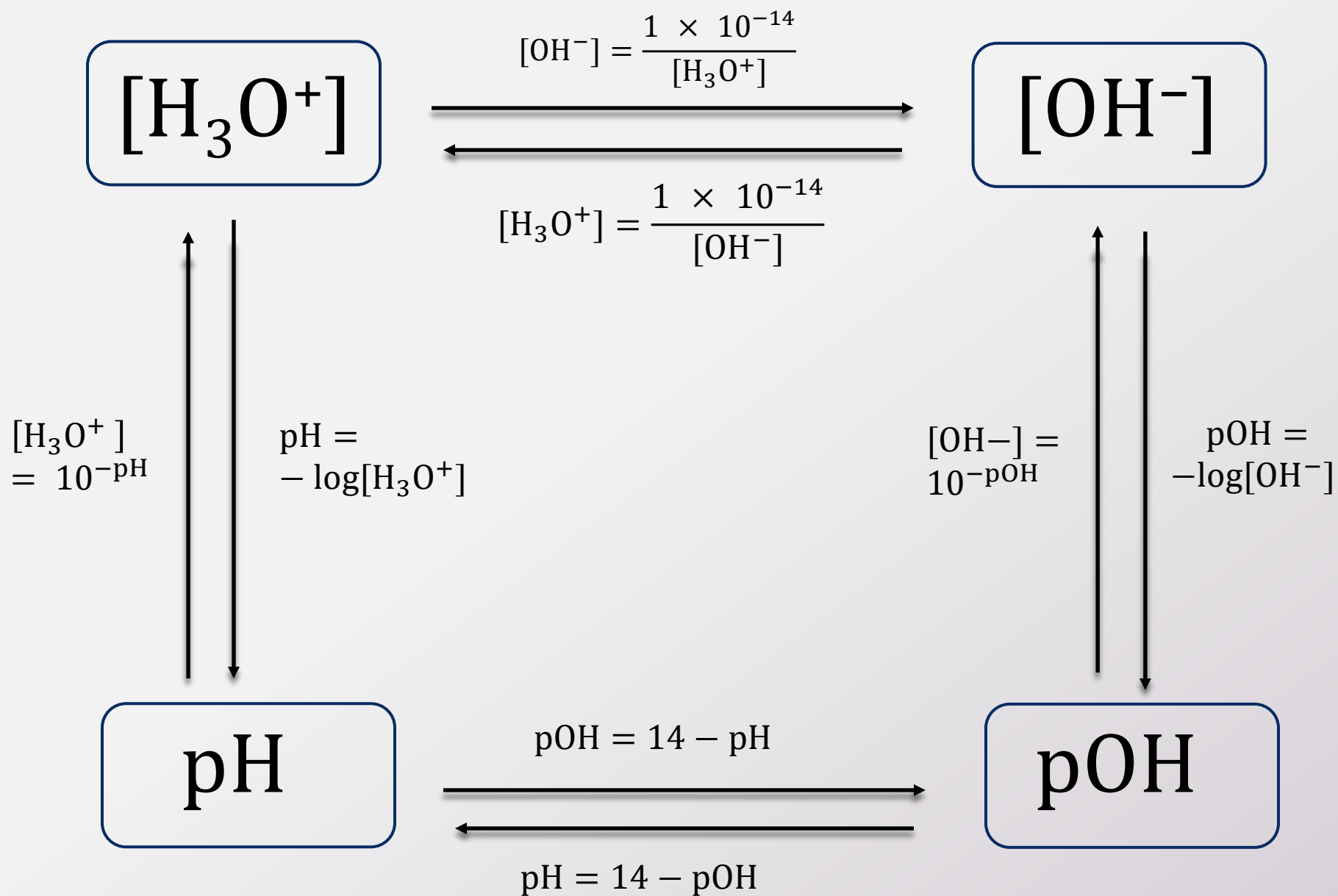
$$[OH^-] = 10^{-pOH}$$

- K_w is the hydronium ion and hydroxide ion concentrations; so pK_w is the addition of pH and pOH.

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

$$pK_w = pH + pOH = 14.00$$

- Next slide has all relationships...



EXAMPLE 3: pOH CALCULATION

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

Answer:

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

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

$$pOH = 1.37$$

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

Answer:

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

$$[H^+] = 10^{-8.35}$$

$$[H^+] = 4.5 \times 10^{-9}$$

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

$$1.0 \times 10^{-14} = 4.5 \times 10^{-9} [OH^-]$$

$$[OH^-] = 1.0 \times 10^{-14} / 4.5 \times 10^{-9} = 2.22 \times 10^{-6} M$$

2nd way:

$$pOH = 14 - 8.35 = 5.65$$

Now use:

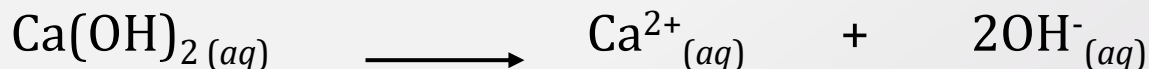
$$[OH^-] = 10^{-pOH}$$

EXAMPLE 4: CALCULATING pH OF BASES

What is the pH of a solution of 0.0340 *M* solution of $\text{Ca}(\text{OH})_2$?

Answer:

$\text{Ca}(\text{OH})_2$ ionizes to give 2 mols of OH^- , so conc. has to be doubled.



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

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

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

$$\text{pOH} = -\log [0.068] = 1.17$$

$$\begin{aligned} \text{pH} &= 14 - \text{pOH} \\ &= 14 - 1.17 = 12.8 \end{aligned}$$

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	$\text{HNO}_3_{(aq)} + \text{H}_2\text{O}_{(l)} \longrightarrow \text{H}_3\text{O}^+_{(aq)} + \text{NO}_3^-_{(aq)}$
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 1 Hydroxides	Group II hydroxides
$\text{LiOH}_{(aq)} \longrightarrow \text{Li}^+_{(aq)} + \text{OH}^-_{(aq)}$	$\text{Ca(OH)}_{2(s)} \longrightarrow \text{Ca}^{2+}_{(aq)} + 2\text{OH}^-_{(aq)}$
$\text{NaOH}_{(aq)} \longrightarrow \text{Na}^+_{(aq)} + \text{OH}^-_{(aq)}$	$\text{Ba(OH)}_{2(s)} \longrightarrow \text{Ba}^{2+}_{(aq)} + 2\text{OH}^-_{(aq)}$

KEY CONCEPTS

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