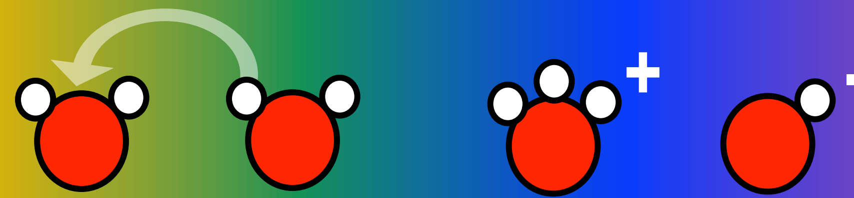
7.8: Aqueous Solutions and the Concept of pH

Remember:

- Have your **7.8 notesheet** ready!
- You can **pause** the video anytime.
- You can **rewind** the video anytime.
- Write down **questions/comments** as you go for discussion in class.

Are you ready???

Part I: The Self-Ionization of Water



$H_2O + H_2O \longleftrightarrow H_3O^+ + OH^-$

water, water, hydronium, hydroxide

- when compounds **dissociate**/ **ionize** in aqueous solution, they produce **ions**.
- these ions are not the only ions in the solution—water itself also **contributes ions** to the **solution—hydronium** (H_3O^+) and **hydroxide** (OH^-)—through self-ionization.
- **self-ionization of water** = process of one water molecule transferring a proton to another water molecule, creating **hydronium & hydroxide**.
  - the **extent** to which water self-ionizes is not very high: In **1 Liter** of pure water, the **concentrations** of H_3O^+ and OH^- are only **1.0×10^{-7} M** for each.
  - these concentrations can be expressed more easily as **$[H_3O^+] = 1.0 \times 10^{-7}$ M** and **$[OH^-] = 1.0 \times 10^{-7}$ M.** The **brackets** indicate "**concentration of**" in **Molarity (M)** *(not molality, m)*.

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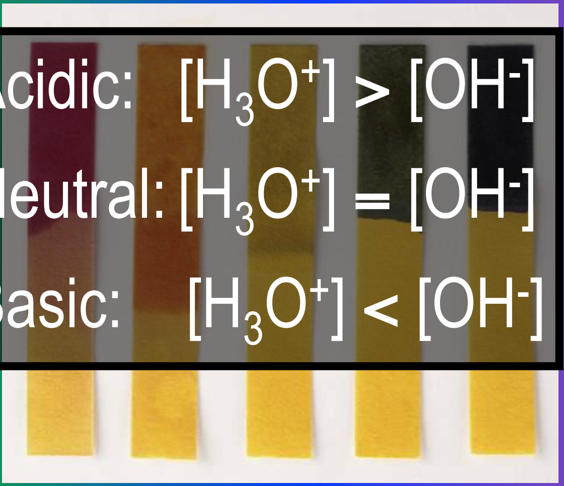
- the **product** of $[H_3O^+]$ and $[OH^-]$ is equal to **1.0 x 10^{-14} M^2**, which is a constant known as the **ionization constant of water, K_W**.

$K_W = [H_3O^+][OH^-] =$ **1.0 x 10^{-14} M^2** at 25°C

Part II: Concentration of Acidic and Basic Solutions

- because the $[H_3O^+]$ and $[OH^-]$ in pure water are **equal**, pure water is **neutral**.
- any substance whose $[H_3O^+]$ **does not equal** its $[OH^-]$ is going to be acidic or basic.
  - if $[H_3O^+]$ is **greater** than 1.0 x 10^{-7} M, the solution is **acidic**.
    - **Ex**: $[H_3O^+]$ of lemon juice = 3.9 x 10^{-3} M

- any substance whose $[H_3O^+]$ **does not equal** its $[OH^-]$ is going to be acidic or basic.
  - if $[H_3O^+]$ is **greater** than 1.0×10^{-7} M, the solution is **acidic**.
    - **<u>Ex</u>**: $[H_3O^+]$ of lemon juice = 3.9×10^{-3} M
  - if $[H_3O^+]$ is less than 1.0×10^{-7} M, the solution
    - **<u>Ex</u>**: $[H_3O^+]$ of Maalox = 3.16×10^{-11} M



Acidic: $[H_3O^+] > [OH^-]$

Neutral: $[H_3O^+] = [OH^-]$

Basic: $[H_3O^+] < [OH^-]$

- to calculate $[H_3O^+]$, you must know $[OH^-]$, and vice versa. Then use the K_W equation rearranged for what you want:
  - **<u>Ex1</u>**: What is the $[H_3O^+]$ if $[OH^-] = 4.56 \times 10^{-8}$ M?

$$[H_3O^+] = \frac{K_W}{[OH^-]}$$

$$[OH^-] = \frac{K_W}{[H_3O^+]}$$

K_W = $\mathbf{1.0 \times 10^{-14}\ M^2}$

$[H_3O^+]$ = ?

$[OH^-] = 4.56 \times 10^{-8}$ M

- to calculate $[H_3O^+]$, you must know $[OH^-]$, and vice versa. Then use the K_W equation rearranged for what you want:
  - **Ex1**: What is the $[H_3O^+]$ if $[OH^-] = 4.56 \times 10^{-8}$ M?

$[H_3O^+] = ?$
$[OH^-] = 4.56 \times 10^{-8}$ M

$$[H_3O^+] = \frac{K_W}{[OH^-]} = \frac{1.0 \times 10^{-14}\ M^2}{4.56 \times 10^{-8}\ M} = \boxed{\mathbf{2.19 \times 10^{-7}\ M}}$$

  - **generally**, solutions whose M of $[H_3O^+]$ has an exponent between -1 and -7 is **acidic**, and solutions with a M of $[H_3O^+]$ with an exponent between -8 and -14 are **basic**.

Part III: The pH/pOH Scale and pH/pOH Calculations

- since the values for $[H_3O^+]$ and $[OH^-]$ are often very small, we have to use scientific notation numbers to express their size. This can be very cumbersome in some calculations.
- a more convenient way of expressing the acidity (or basicness) of a solution is **pH** and **pOH**.

- since the values for $[H_3O^+]$ and $[OH^-]$ are often very small, we have to use scientific notation numbers to express their size. This can be very cumbersome in some calculations.
- a more convenient way of expressing the acidity (or basiciness) of a solution is **pH** and **pOH**.
  - **pH** = the negative logarithm of the **hydronium** ion concentration
  - **pOH =** the negative logarithm of the **hydroxide** ion concentration
- these values allow a number in scientific notation to be put into regular notation.

$$\boxed{pH = -\log[H_3O^+]}$$

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

Ex2: What is the pOH if $[OH^-] = 7.13 \times 10^{-9}$ M?

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pOH= ?
$[OH^-] = 7.13 \times 10^{-9}$ M

$pOH = -\log[OH^-] = -\log(7.13 \times 10^{-9}\ M) = \boxed{\mathbf{8.15}}$

- since values of $[H_3O^+]$ and $[OH^-]$ are related by K_W (whose $-\log = 14.0$), the following relationship exists:

pH + pOH = 14.0

- the values on the pH/pOH scale are as follows:

pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14	pH
$[H_3O^+]$	1	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}	$[H_3O^+]$
$[OH^-]$	10^{-14}	10^{-13}	10^{-12}	10^{-11}	10^{-10}	10^{-9}	10^{-8}	10^{-7}	10^{-6}	10^{-5}	10^{-4}	10^{-3}	10^{-2}	10^{-1}	1	$[OH^-]$
pOH	14	13	12	11	10	9	8	7	6	5	4	3	2	1	0	pOH

acidic values　　neutral　　basic values

- notice that **pH** values from 0 to 6.9 are **acidic**, 7.1 to 14 are **basic**, and 7.0 is **neutral**. In the same respect, **pOH** values from 0 to 6.9 are **basic**, 7.1 to 14 are **acidic**, and 7.0 is **neutral**.

pH	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14	pH
$[H_3O^+]$	1	10^{-1}	10^{-2}	10^{-3}	10^{-4}	10^{-5}	10^{-6}	10^{-7}	10^{-8}	10^{-9}	10^{-10}	10^{-11}	10^{-12}	10^{-13}	10^{-14}	$[H_3O^+]$
$[OH^-]$	10^{-14}	10^{-13}	10^{-12}	10^{-11}	10^{-10}	10^{-9}	10^{-8}	10^{-7}	10^{-6}	10^{-5}	10^{-4}	10^{-3}	10^{-2}	10^{-1}	1	$[OH^-]$
pOH	14	13	12	11	10	9	8	7	6	5	4	3	2	1	0	pOH

acidic values neutral basic values

- notice that **pH** values from 0 to 6.9 are acidic, 7.1 to 14 are basic, and 7.0 is **neutral**. In the same respect, **pOH** values from 0 to 6.9 are basic, 7.1 to 14 are acidic, and 7.0 is neutral.
- the pH values of some common substance are listed in the table to the right:
- let's try two more example problems:

Material	pH	Material	pH
Gastric juice	1.0–3.0	Saliva	6.5–7.5
Lemons	2.2–2.4	Pure water	7.0
Vinegar	2.4–3.4	Blood	7.3–7.5
Soft drinks	2.0–4.0	Eggs	7.6–8.0
Oranges	3.0–4.0	Sea water	8.0–8.5
Tomatoes	4.0–4.4	Antacids	8.7
Bread	5.0–6.0	Baking soda	9.5
Rainwater	5.4–5.8	Milk of magnesia	10.5
Potatoes	5.6–6.0	Ammonia	12.3
Milk	6.3–6.6	Drain cleaner	12.7-13.8

- let's try two more example problems:
  - **Ex3**: What is the pH if $[OH^-] = 1.45 \times 10^{-3}$ M?

pH = ?

$[OH^-] = 1.45 \times 10^{-3}$ M

$pOH = -\log[OH^-] = -\log(1.45 \times 10^{-3}\ M) = 2.84$

$pH + pOH = 14$ $pH = 14 - pOH$ $pH = 14 - 2.84 = \boxed{\mathbf{11.16}}$

  - **Ex4**: What is the pOH if $[H_3O^+] = 8.05 \times 10^{-12}$ M?

pOH = ?

$[H_3O^+] = 8.05 \times 10^{-12}$ M

$pH = -\log[H_3O^+] = -\log(8.05 \times 10^{-12}\ M) = 11.09$

$pH + pOH = 14$ $pOH = 14 - pH$ $pOH = 14 - 11.09 = \boxed{\mathbf{2.91}}$

Part IV: Converting pH/pOH Values into $[H_3O^+]$ and $[OH^-]$ Values

- if you want to know the $[H_3O^+]$ or $[OH^-]$ from a given pH or pOH value, use the following equations: $[H_3O^+] = 10^{-pH}$ $[OH^-] = 10^{-pOH}$

Part IV: Converting pH/pOH Values into $[H_3O^+]$ and $[OH^-]$ Values

- if you want to know the **$[H_3O^+]$** or **$[OH^-]$** from a given pH or pOH value, use the following equations:

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

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

- **Ex5**: What is the $[OH^-]$ if the pH = 12.27 ?

pH = 12.27
pOH = ?
$[OH^-]$ = ?

$pOH = 14 - pH$ $\quad pOH = 14 - 12.27 = 1.73$

$[OH^-] = 10^{-pOH}$ $\quad [OH^-] = 10^{-1.73} = 0.0186 = \boxed{\mathbf{1.86 \times 10^{-2}\ M}}$

- **Ex6**: What is the $[H_3O^+]$ if pH = 5.16 ?

$[H_3O^+]$ = ?
pH = 5.16

$[H_3O^+] = 10^{-pH}$ $\quad [H_3O^+] = 10^{-5.16} = \boxed{\mathbf{6.92 \times 10^{-6}\ M}}$

- Make sure notesheet is **completely filled in**
- Preview the **funsheets** (7.8a, b, c)
- **Rewind and review** any parts that were not clear
- Bring both **notesheet** and **funsheet packets** to class