Chemistry 30 Unit 3 - Acid-Base Chemistry

The Water Ionization Constant, K_w

Today, you will...

- define Kc to predict the extent of the reaction and write equilibrium-law expressions for given chemical equations, using lowest whole-number coefficients

Diploma Alert!

20. If the $[H^+_{(aq)}]$ in reaction I is 0.020 mol/L, then the pH and pOH are, respectively,

A. 1.05 and 12.95
B. 1.40 and 12.60
C. 1.70 and 12.30
D. 2.00 and 12.00

Ans:

Diploma Alert!

Numerical Response

10. The concentration of $H_3O^+_{(aq)}$ ions in a particular bottle of wine is 3.2×10^{-4} mol/L. The pH of this wine is __________.

(Record your **three-digit answer** in the numerical-response section on the answer sheet.)

Ans: [redacted]

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Numerical Response

12. If the pH of a sample of rainwater is 3.2, then the pOH is __________.

(Record your **three-digit answer** in the numerical-response section on the answer sheet.)

Ans: [redacted]

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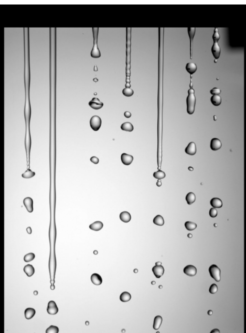
31. Sour pickles have a pH of about 3.00. The $[OH^-_{(aq)}]$ in a typical sour pickle is

A. 1.0×10^{-11} mol/L
B. 3.0×10^{-11} mol/L
C. 1.0×10^{-3} mol/L
D. 3.0×10^{-3} mol/L

Ans: [redacted]



Chemistry = Lies

Misconceptions (well, sort of...)

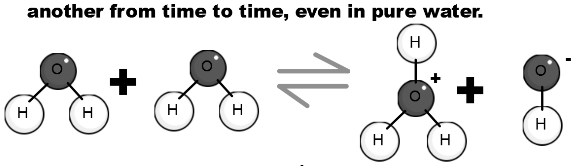
Common knowledge says that water can conduct electricity (i.e. don't drop a toaster in your bathtub...), but we later said that this was because of the impurities in tap water.

It was commonly agreed on that pure water doesn't conduct electricity because water is not an electrolyte (i.e. does not form charged particles).

Unfortunately, this is only part of the story...

Recall our favorite chemistry theory, the <u>Particle Theory</u>!

Particle theory says that all particles are constantly moving and bumping into one another. Therefore, water molecules must bump into one another from time to time, even in pure water.



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

Equilbrium theory also says that at least a small amount of water might dissociate and form hydroxide and hydronium ions.

These ions cause even <u>pure</u> water to have a very slight conductivity!

So where does the equilibrium lie in this reaction? Just how many ions are produced?

To determine this, we could apply the equilibrium law to this equation:

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

$$K_w = \frac{[H_3O^+][OH^-]}{[HOH]^2}$$

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

***as water is a pure liquid and not included...**

$$\boxed{K_w = [H_3O^+][OH^-]}$$

This is known as the water equilibrium expression. After careful measurement, chemists have established the value of K_w to be 1.00×10^{-14} at SATP.

Notice: hydronium and hydroxide are produced in equal amounts in pure water. As pure water is pH neutral, <u>the concentrations of hydronium and hydroxide must be equal in any neutral solution.</u>

Important Note:

$K_w = 1.00 \times 10^{-14}$ only at SATP. At other temperatures, the value will vary:

$K_w = 6.76 \times 10^{-15}$ at 20°C
$K_w = 1.47 \times 10^{-14}$ at 30°C

$$[H_3O^+][OH^-] = 1.00 \times 10^{-14} \text{ mol/L}$$

$$[H_3O^+] = [OH^-] = \sqrt{(1.00 \times 10^{-14})} = 1.00 \times 10^{-7} \text{ mol/L}$$

As the concentrations of hydroxide and hydronium are equal, we can determine the concentrations of each in a neutral solution by square rooting the K_w constant (assume units of mol/L).

So in a neutral solution:

$[H_3O^+] = 1.00 \times 10^{-7}$ mol/L
$[OH^-] = 1.00 \times 10^{-7}$ mol/L

According to Arrhenius, an acid has a high concentration of hydronium ions, and a base will have a higher concentration of hydroxide ions.

But the water equilibrium expression will always hold, for acidic or basic solutions, so long as they are dissolved in water.

This provides us with a nice way to determine the concentration of either hydronium or hydroxide in aqueous solutions.

ex) A 0.15 mol/L solution of $HCl_{(aq)}$ at SATP is found to have a hydronium concentration of 0.15 mol/L. What is the $[OH^-]$ in this solution?

$[H_3O^+][OH^-] = 1.00 \times 10^{-14}$ mol/L

therefore;

$$[OH^-] = \frac{1.00 \times 10^{-14} \text{ mol/L}}{[H_3O^+]}$$

$$[OH^-] = \frac{1.00 \times 10^{-14} \text{ mol/L}}{[0.15]} = 6.7 \times 10^{-14} \text{ mol/L}$$

ex) Calculate the $[H_3O^+]$ in a 0.25 mol/L solution of barium hydroxide.

*Hint: You must use stoich!

ex) Determine the $[H_3O^+]$ and $[OH^-]$ in 500 mL of an aqueous solution for home-made soap containing 2.6 g of sodium hydroxide.

Relationships between pH and pOH

Recall from water equilibrium that:

$$1.00 \times 10^{-14} = [H_3O][OH^-]$$

If we were to take the log of the left and right side of this equation:

$$\log(1.00 \times 10^{-14}) = \log[H_3O] + \log[OH^-]$$

$$-14 = -pH + -pOH$$

$$\boxed{14 = pH + pOH}$$

This shows the relationship between pH and pOH. In any given solution, they must add up to 14.

On to Brosted-Lowery!