c. 6.00×10^{-3}M $Sr(OH)_2$

5. Hebden p. 127 #28, 29

XII) pH

What does pH stand for?

"power of H", meaning the exponent (logarithm) of the H+ molarity.

pH is an indication of the acidity/basicity of a solution.

pH is the negative logarithm of $[H^+]$ or $[H_3O^+]$ in a solution: $-\log[H^+]$ or $-\log[H_3O^+]$

For example, take a solution with $[H_3O^+] = 1.0 \times 10^{-7}$M. The log is -7, so negative log (the pH) is –(-7), or 7.

What kind of solution has $[H_3O^+] = 1.0 \times 10^{-7}$M?

pH = 7 is neutral, such as pure water

What is the pH of a solution that has $[H_3O^+] = 1.0 \times 10^{-4}$M, and is the solution acidic, basic, or neutral?

$pH = -\log(1.0 \times 10^{-4} M) = 4.00$ solution is acidic $[H_3O^+] > [OH^-]$

$[OH^-]$ would $= 1.0 \times 10^{-10}$ M

What is the pH of a solution that has $[H_3O^+] = 1.0 \times 10^{-11}$M and is the solution acidic, basic, or neutral? $pH = -\log(1.0 \times 10^{-11}) = 11.00$ $[H_3O^+] < [OH^-]$ BASIC

* The $\boxed{\text{pH scale}}$ is generally considered to be from 0 to 14 at 25°C.

pH values can sometimes be below 0 (very acidic solutions) or above 14 (very basic. solutions).

animation.

Complete the table below (Remember: $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$)

$[H_3O^+]$ (M)	1.0×10^1	1×10^0	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}	10^{-15}
pH	-1	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
$[OH^-]$	1.0×10^{-15}	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}	10^{0}	10^{1}
acidic, basic, or neutral?	A	A	A	A	A	A	A	A	N	B	B	B	B	B	B	B	B

Remember that $[H_3O^+]$ and $[OH^-]$ are inversely related (as one goes up, the other goes down). Thus, a high $[H_3O^+]$ in a solution corresponds to a low $[OH^-]$, as their product must always equal the $K_w = 1.0 \times 10^{-14}$ at 25°C.

If pH decreases by 1, what happens to $[H_3O^+]$? $[OH^-]$?

remember -log of a - exponent

pH ↓ by 1 = $[H_3O^+]$ increase by 10 x

= $[OH^-]$ decrease by 10x

If pH increases by 1, what happens to $[H_3O^+]$? $[OH^-]$?

pH ↑ 1 = $[H_3O^+]$ decreases by 10x
$[OH^-]$ increases by 10x

What is the pH of a 1.0×10^{-6}M H_3O^+ solution?

↳ pH = 6.00 $(-\log(1.0\times10^{-6}) = 6.00)$

Is the pH of a 4.2×10^{-6}M H_3O^+ solution greater than 6 or less than 6? How do you know?

pH: 4.2×10^{-6}M > 1.0×10^{-6}M ∴ more $[H_3O^+]$ ∴ more acidic
This means the pH must be less than 6.

pH is defined as $-\log [H_3O^+]$, and can be found using a calculator:

press (-) … then log … then 4.2 EXP/EE/x10^-6 …. then "="
(or 4.2 EXP-6, then log … then you change the sign +/-) pH=5.38

How do 'sig figs' work when calculating pH?

$\underbrace{4.2}_{2\text{ s.f.}}\times10^{-6}$ M ← how ever many sig figs in Molarity, that is how many digits go AFTER decimal in pH.

Find the pH of each solution below with proper sig figs:

$[H_3O^+]$(M)	2.15×10^{-2}	8×10^{-9}	9.334×10^{-5}	5.0×10^{-13}	3.500
pH	1.668	8.1	4.0299	12.30	-0.5441
A, B, or N	A	B	A	B	A

$[H_3O^+]$ is calculated from pH by the following:

$[H_3O^+] = 2^{nd} \log(-pH)$ *2^{nd} same as *shift* or *inv* on calc (or the 10^x button.)

don't forget the (-) sign!!

Find $[H_3O^+]$ for each with proper sig figs:

$[H_3O^+]$ (M)	4.78×10^{-8}	2.8×10^{-5}	5×10^{-2}	6.0×10^{-14}	2.368×10^{-16}
pH	7.321	4.56	1.3	13.22	15.6257
A, B, or N	B	A	A	B	B

number of sig figs needed in [conc.]

$10^{\wedge}(-7.321) = 4.77529\times10^{-8}$ M

pH has 3 d.p. → $[H_3O^+]$ needs 3 sig figs *pH + pOH = 14* always

pOH

What is pOH?

"power of hydroxide" OH^-

How do you calculate it if you know $[OH^-]$?

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

How do you calculate $[OH^-]$ if you know the pOH?

2^{nd}/inv log(-pOH) or 10^(-pOH) } depending on calc. buttons.

If $[H_3O^+] = 3.45 \times 10^{-5}$M, find pH, $[OH^-]$, and pOH. Is solution A, B, or N?

① $pH = -\log(3.45 \times 10^{-5} M)$

$pH = 4.462$ s.f.

② $K_w = [H_3O^+][OH^-]$

$\therefore [OH^-] = \frac{1.0 \times 10^{-14}}{4.462M}$

$[OH^-] = 2.9 \times 10^{-10} M$

③ $pOH = -\log(2.9 \times 10^{-10} M)$

$\boxed{pOH = 9.54 \quad \text{acidic}}$

If $[OH^-] = 7.2 \times 10^{-3}$M, find pOH, $[H_3O^+]$, and pH. Is solution A, B, or N?

① $pOH = -\log(7.2 \times 10^{-3} M)$

$pOH = 2.14$ s.f.

② $K_w = [H_3O^+][OH^-]$

$\therefore [H_3O^+] = \frac{1.0 \times 10^{-14}}{7.2 \times 10^{-3} M}$

$[H_3O^+] = 1.4 \times 10^{-12} M$

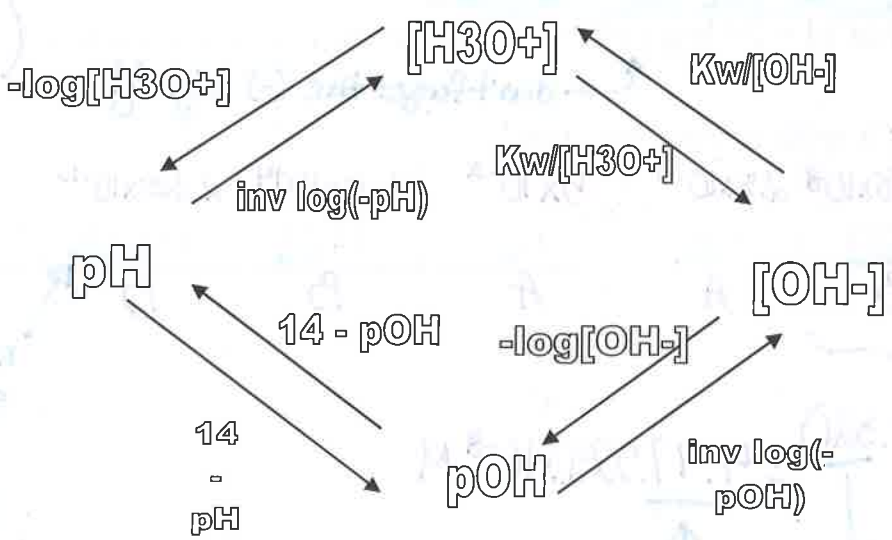
③ $pH = -\log(1.4 \times 10^{-12} M)$

$\boxed{pH = 11.86 \quad \text{basic}}$

Using the results of the last two examples what relationship exists between pH and pOH at 25°C?

pH + pOH = 14

Therefore if pH < 7, then pOH __> 7__ and the solution is __ACIDIC__.
If pH > 7, then pOH __< 7__ and the solution is __BASIC__.



known: $pOH = 8.421$ $\therefore pH = 5.579$ $(14 - 8.421)$

Example: Find the pH, $[H_3O^+]$, and $[OH^-]$ of a solution that has a pOH of 8.421. $[OH^-] = 10^{\wedge}(-8.421)$ $[H_3O^+] = 10^{\wedge}(-5.579)$

$[OH^-] = 3.79 \times 10^{-9} M$ $[H_3O^+] = 2.64 \times 10^{-6} M$

Assignment 8: pH/pOH Exercises

*Quizizz for practice *

1. Find the pH, pOH, and $[OH^-]$ if
   a) $1.0 \times 10^{-5}M\ H_3O^+$.
   b) $2.65 \times 10^{-7}M\ H_3O^+$.
   c) $6.744 \times 10^{-12}M\ H_3O^+$.
2. Find the $[H_3O^+]$, pOH, and $[OH^-]$ if
   a) pH = 2.35
   b) pH = 6.456
   c) pH = 10.76
3. Find the $[OH^-]$, $[H_3O^+]$, and pH if
   a) pOH = 2.34
   b) pOH = 12.59
   c) pOH = 7.10
4. Hebden p.141 #55, 56

pK_w

The 'p' of any value is the –log of that value. For example, the pH of $[H_3O^+]$ is $-\log[H_3O^+]$, and the pOH of $[OH^-]$ is $-\log[OH^-]$. Therefore, how would you calculate the pK_w at 25°C if $K_w = 1.0 \times 10^{-14}$?

$pK_w = -\log K_w = -\log(1.0 \times 10^{-14}) = 14.00$

sig. figs

How do pH and pOH relate to each other?

pH and pOH $pH + pOH = 14.00$ $\therefore pH + pOH = pK_w$

This is because when you multiply powers in math, the shortcut is to add their exponents!