

Announcements

Discussion assignment 2 replies due Wed before class.

Ch 14 MC assignment is due Wed before class.

Extra credit labs: experiments 11 and 12 in your packet. Do them at home and bring the completed worksheets to the final exam. They are worth up to 5 extra credit points each.

Comprehensive final exam is this Wednesday, May 13 from **10 am to noon**. The exam is 50 3-point multiple choice questions. (150 points total).

You are allowed **one handwritten 3x5" note card** (both sides) for the final exam.

Review sessions Monday at 3pm, Tuesday at 1pm and Wednesday at 8am.

pH values and calculations

pH = 2 means $[H_3O^+] = 10^{-2} M$

$.01 M$

pH = 3 means $[H_3O^+] = 10^{-3} M$

$.001 M$

$10^0 = 1$

$10^{-1} = \frac{1}{10} = .1$

$10^{-2} = \frac{1}{100} = .01$

$10^{-3} = \frac{1}{1000} = .001$

Which is **more acidic**, pH = 2 or pH = 3?

$[H_3O^+] = .01 M$

$.001 M$

0

1 $[H_3O^+]$

2 $\rightarrow 10^{-2} M$

3 $\rightarrow 10^{-3} M$

4

5

6

(neut) 7 $\rightarrow 10^{-7} M$

8

9 $\rightarrow 10^{-9} M$

10

11

12

13

14

If $[H_3O^+] = 10^{-9} M$, then pH = 9

Logarithms:

$\log 10^9 = 9$

$\log 10^2 = 2$

$\log 10^{-9} = -9$

$-\log 10^{-9} = 9$

$.007$

$.008$

$.009$

$.01$

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

exact

exact

If $[H_3O^+] = 10^{-6} M$, then pH = $-\log 10^{-6} = 6$

2 sf

2 dp

If $[H_3O^+] = 1.0 \times 10^{-12} M$, pH = 12.00

Sig figs in concentrations become decimal places in pH

If $[H_3O^+] = 4.2 \times 10^{-12} M$, pH = $-\log (4.2 \times 10^{-12}) = 11.38$

EE, EXP

Text calculator: $(-)$ LOG 4.2 E $(-)$ 12 =

Numerical calculator: 4.2 E 12 +/- LOG +/-

pH calculations

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

If $[\text{H}_3\text{O}^+] = 2.87 \times 10^{-3} \text{ M}$, what is pH?

$$\text{pH} = -\log(2.87 \times 10^{-3}) = \underline{2.542}$$

(in pH values, the sig figs are after the decimal point)

If $\text{pH} = 8$, $[\text{H}_3\text{O}^+] = 10^{-8} \text{ M}$

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

If $\text{pH} = 2.87$, $[\text{H}_3\text{O}^+] = 10^{-2.87} = .00135 \rightarrow .0013 \text{ M}$
need exact power of 10
2 s.f.
 $\boxed{= 1.3 \times 10^{-3} \text{ M}}$

text calculator: 10^x $(-)$ 2.87 $=$

numerical calculator: 2.87 $+/-$ 10^x

If $\text{pH} = 6.43$, $[\text{H}_3\text{O}^+] = 10^{-6.43} = 3.7 \times 10^{-7} \text{ M}$

(Use MODE or SCI if your calculator gives you 0.00000...)

[OH⁻] and pOH calculations

If [OH⁻] = 10⁻⁴ M, pOH = 4

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

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

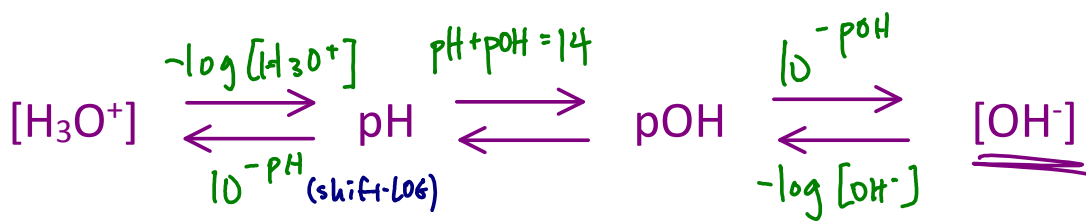
$$\text{pH} + \text{pOH} = 14$$

pH = 12.23

pOH = 14 - pH

= 1.77

pH		pOH
0	acidic	14
1		13
2		12
3		11
4		10
5		9
6		8
7	neut.	7
8		6
9		5
10		4
11		3
12		2
13		1
14	basic	0



If [OH⁻] = 0.0230 M, pOH = -log(.0230) = 1.638

pH = 14 - 1.638 = 12.362

[H₃O⁺] = 10^{-12.362} = 4.35 × 10⁻¹³ M
(shift-log)

pH shows this solution is: acidic / basic / neutral

Calculations for specific solutions

What is the pH, pOH, $[\text{OH}^-]$, and $[\text{H}_3\text{O}^+]$ of $1.50 \times 10^{-3} \text{ M HCl(aq)}$?

Which is HCl? acid or base

$1 \text{ mol HCl} \rightarrow 1 \text{ mol H}_3\text{O}^+$ in water

$$[\text{HCl}] = 1.50 \times 10^{-3} \text{ M}$$

$$[\text{H}_3\text{O}^+] = 1.50 \times 10^{-3} \text{ M}$$

$$\begin{aligned} \text{pH} &= -\log(1.50 \times 10^{-3}) \\ &= 2.824 \text{ (acidic)} \end{aligned}$$

$$\text{pOH} = 14 - 2.824 = 11.176$$

$$[\text{OH}^-] = 10^{-11.176} = 6.67 \times 10^{-12} \text{ M}$$

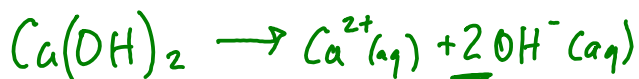
These equations will be given on the final exam:

- $\text{pH} = -\log [\text{H}_3\text{O}^+]$
- $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$
- $\text{pH} + \text{pOH} = 14$
- $\text{pOH} = -\log [\text{OH}^-]$
- $[\text{OH}^-] = 10^{-\text{pOH}}$

What is the pH, pOH, $[\text{OH}^-]$, and $[\text{H}_3\text{O}^+]$ of $2.43 \times 10^{-5} \text{ M KOH(aq)}$?

Which is KOH? acid or base

$1.35 \times 10^{-2} \text{ M}$ calcium hydroxide soln



$$[\text{Ca(OH)}_2] = 1.35 \times 10^{-2} \text{ M}$$

$$[\text{OH}^-] = 1.35 \times 10^{-2} \text{ M} \times 2 = 2.70 \times 10^{-2} \text{ M}$$

$$\text{pOH} = 1.569$$

$$\text{pH} = 12.431 \text{ (basic)}$$

$$[\text{H}_3\text{O}^+] = 3.71 \times 10^{-13} \text{ M}$$

$$[\text{KOH}] = 2.43 \times 10^{-5} \text{ M}$$

$$[\text{OH}^-] = 2.43 \times 10^{-5} \text{ M}$$

$$\begin{aligned} \text{pOH} &= -\log(2.43 \times 10^{-5}) \\ &= 4.614 \end{aligned}$$

$$\text{pH} = 14 - 4.614 = \underline{9.386}$$

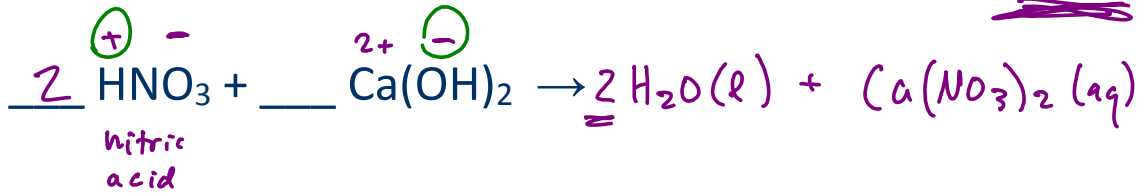
(solution is basic)

$$\begin{aligned} [\text{H}_3\text{O}^+] &= 10^{-9.386} \\ &= 4.11 \times 10^{-10} \text{ M} \\ &\text{or } 4.12 \times 10^{-10} \text{ M} \end{aligned}$$

depending on using intermediate values

Acid-base neutralization reactions

When an acid and a base react, treat them like ionic compounds and do a double displacement reaction, remembering that H^+ and OH^- combine to form: $H_2O(l)$



phosphoric acid + magnesium hydroxide \rightarrow

$$\underline{2} \overset{+}{H_3} \overset{3-}{PO_4} + \underline{3} \overset{2+}{Mg} \overset{-}{(OH)_2} \rightarrow \underline{6} H_2O(l) + \underline{\quad} Mg_3(PO_4)_2$$

How many L of 0.25M MOL HCl will completely neutralize 2.61 g of aluminium hydroxide?



$$2.61 \text{ g } Al(OH)_3 \times \frac{1 \text{ mol } Al(OH)_3}{43.99 \text{ g } Al(OH)_3} \times \frac{3 \text{ mol HCl}}{1 \text{ mol } Al(OH)_3} \times \frac{1 \text{ L HCl soln}}{0.25 \text{ mol HCl}} = \boxed{0.71 \text{ L HCl soln}}$$