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 → 10^{-2} M
3 → 10^{-3} M
4
5
6
(neut) 7 → 10^{-7} M
8
9 → 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

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

.007
.008
.009
.01

If $[H_3O^+] = \underbrace{10^{-6}}_{\text{exact}}$ *M*, then pH = -log 10^{-6} = $\underbrace{6}_{\text{exact}}$

If $[H_3O^+] = 1.0 \times 10^{-12}$ *M*, pH = 12.00   (2 sf → 2 dp)

sig figs in concentrations become decimal places in pH

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

Text calculator: $\boxed{(-)}$ $\boxed{\text{LOG}}$ 4.2 $\boxed{\text{E}}$ (EE, Exp) $\boxed{(-)}$ 12 $\boxed{=}$

Numerical calculator: 4.2 $\boxed{\text{E}}$ 12 $\boxed{+/-}$ $\boxed{\text{LOG}}$ $\boxed{+/-}$

pH calculations

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

If $[H_3O^+] = 2.87 \times 10^{-3}$ *M*, what is pH?

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

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

If pH = 8, $[H_3O^+] =$ 10^{-8} M

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

need exact power of 10

2 dp

2 s.f.

If pH = 2.87, $[H_3O^+] = 10^{-2.87} =$ .00135 $\rightarrow$ .0013 M $\boxed{= 1.3 \times 10^{-3} \text{ M}}$

shift LOG

text calculator: $\boxed{10^x}$ $\boxed{(-)}$ 2.87 $\boxed{=}$

shift LOG

numerical calculator: 2.87 $\boxed{+/-}$ $\boxed{10^x}$

If pH = 6.43, $[H_3O^+] = 10^{-6.43} =$ 3.7×10^{-7} M

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

$[OH^-]$ and pOH calculations

If $[OH^-] = 10^{-4}$ *M*, pOH = 4

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

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

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

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

$$[H_3O^+] \underset{10^{-pH}\text{ (shift-LOG)}}{\overset{-\log[H_3O^+]}{\rightleftarrows}} pH \overset{pH+pOH=14}{\rightleftarrows} pOH \underset{-\log[OH^-]}{\overset{10^{-pOH}}{\rightleftarrows}} [OH^-]$$

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

pH = 14 − 1.638 = 12.362

$[H_3O^+] = 10^{-12.362}$ (shift-LOG) = 4.35×10^{-13} M

pH shows this solution is: acidic / basic / neutral

Calculations for specific solutions

What is the pH, pOH, $[OH^-]$, and $[H_3O^+]$ of
1.50×10^{-3} *M* HCl(*aq*)?

Which is HCl? acid or base

1 mol HCl → 1 mol H_3O^+ in water

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

$[H_3O^+] = 1.50 \times 10^{-3}$ M

$pH = -\log(1.50 \times 10^{-3})$

$= 2.824$ (acidic)

$pOH = 14 - 2.824 = 11.176$

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

These equations will be given on the final exam:

- $pH = -\log[H_3O^+]$
- $[H_3O^+] = 10^{-pH}$
- $pH + pOH = 14$
- $pOH = -\log[OH^-]$
- $[OH^-] = 10^{-pOH}$

What is the pH, pOH, $[OH^-]$, and $[H_3O^+]$ of
2.43×10^{-5} *M* KOH(*aq*)?

Which is KOH? acid or base

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

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

$pOH = -\log(2.43 \times 10^{-5})$

$= 4.614$

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

(solution is basic)

$[H_3O^+] = 10^{-9.386}$

$= 4.11 \times 10^{-10}$ M

or 4.12×10^{-10} M

depending on using intermediate values

1.35×10^{-2} M calcium hydroxide soln

$Ca(OH)_2 \rightarrow Ca^{2+}(aq) + \underline{2}\,OH^-(aq)$

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

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

$pOH = 1.569$

$pH = 12.431$ (basic)

$[H_3O^+] = 3.71 \times 10^{-13}$ M

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: $H_2O\,(l)$

$\underline{2}\ HNO_3 + \underline{\quad}\ Ca(OH)_2 \rightarrow \underline{2}\ H_2O(l) + Ca(NO_3)_2\,(aq)$

nitric acid

$\underline{\quad}\ H_2SO_4 + \underline{2}\ KOH \longrightarrow \underline{2}\ H_2O\,(l) + \underline{\quad}\ K_2SO_4\,(aq)$

phosphoric acid + magnesium hydroxide →

$\underline{2}\ H_3PO_4 + \underline{3}\ Mg(OH)_2 \longrightarrow \underline{6}\ H_2O\,(l) + \underline{\quad}\ Mg_3(PO_4)_2$

How many L of 0.25 M (mol/L) HCl will completely neutralize 2.61 g of aluminum hydroxide?

$\underline{3}\ HCl + Al(OH)_3 \longrightarrow \underline{3}\ H_2O + AlCl_3$

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