

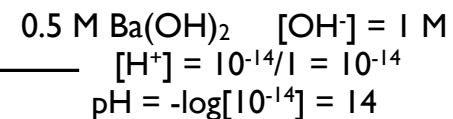
Exam
Calculators Yes!

More time for you to finish

Yes
7:30 - 9:30

What is the pH of a 0.5M solution of barium hydroxide?

- A. 0.5
- B. 9
- C. 14
- D. 12
- E. 10



What is the pOH of a 0.5M solution of barium hydroxide?

- A. 0
 - B. 0.5
 - C. 1
 - D. 12
 - E. 14
- 0.5 M Ba(OH)₂ [OH⁻] = 1 M
- pOH = -log[1] = 0
- pOH = 14 - pH

Which of the following is the most acidic?

- A. acetic acid $K_a = 1.8 \times 10^{-5}$
- B. hydrofluoric acid $K_a = 7.2 \times 10^{-4}$
- C. hydrocyanic acid $K_a = 6.2 \times 10^{-10}$
- D. nitrous acid $K_a = 4.0 \times 10^{-4}$

Which of the following is the most basic?

- A. ammonia $K_b = 1.8 \times 10^{-5}$
- B. methyl amine $K_b = 4.38 \times 10^{-4}$
- C. ethyl amine $K_b = 5.6 \times 10^{-4}$ ← largest K_b
- D. pyridine $K_b = 1.7 \times 10^{-9}$

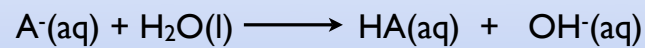
Which of the following is the most basic?

- A. acetate acetic acid $K_a = 1.8 \times 10^{-5}$
- B. fluoride hydrofluoric acid $K_a = 7.2 \times 10^{-4}$
- C. cyanide hydrocyanic acid $K_a = 6.2 \times 10^{-10}$ ← smallest K_a
will be largest K_b
- D. nitrite nitrous acid $K_a = 4.0 \times 10^{-4}$

Converting K_a to K_b



$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$



$$K_b = \frac{[\text{OH}^-][\text{HA}]}{[\text{A}^-]}$$

$$K_a \times K_b = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} \times \frac{[\text{OH}^-][\text{HA}]}{[\text{A}^-]} = [\text{H}^+][\text{OH}^-] = K_w$$

K_a for benzoic acid is 6×10^{-5} ,
what is K_b for the benzoate ion?

- A. 6×10^{-5}
- B. 6×10^{-9}
- C. 1.6×10^{-10} ← $K_b = K_w / K_a = 10^{-14} / 6 \times 10^{-5}$
 1.6×10^{-10}
- D. 1.6×10^{-14}
- E. 60×10^{-19}

Which is the most soluble?

- A. Aluminum Hydroxide $\text{Al}(\text{OH})_3$ $K_{\text{sp}} = 3 \times 10^{-34}$
 B. Barium Fluoride BaF_2 $K_{\text{sp}} = 1.8 \times 10^{-7}$ ←
 C. Calcium Sulfate $\text{Ca}(\text{SO}_4)$ $K_{\text{sp}} = 5 \times 10^{-5}$

What is the solubility of ScF_3 ?



$$K_{\text{sp}} = [\text{Sc}^{3+}][\text{F}^{-}]^3 = 4.2 \times 10^{-18}$$

	ScF_3	Sc^{3+}	F^{-}	$K = [\text{Sc}^{3+}][\text{F}^{-}]^3$
I	n_{solid}	0	0	$K = (x)(3x)^3$
C	-x	+x	+3x	$K = 27x^4 = 4.2 \times 10^{-18}$
E	n-x	+x	+3x	$x = 1.99 \times 10^{-5}$

$$x = [\text{Sc}^{3+}]$$

x is also the number of moles of ScF_3 that dissolve
 molar solubility 1.99×10^{-5} moles/L
 solubility 2×10^{-3} g/L

Same solution

When you have only one compound in water
 What you need to know if the “generic formula”

MX	$K_{\text{sp}} = [\text{M}^{+}][\text{X}^{-}] = x^2$	$x = (K_{\text{sp}})^{1/2}$
MX_2	$K_{\text{sp}} = [\text{M}^{+}][\text{X}^{-}]^2 = 4x^3$	$x = (K_{\text{sp}}/4)^{1/3}$
MX_3	$K_{\text{sp}} = [\text{M}^{+}][\text{X}^{-}]^3 = 27x^4$	$x = (K_{\text{sp}}/27)^{1/4}$

Which is the most soluble?

- A. Iron (III) Hydroxide $K_{\text{sp}} = 2.8 \times 10^{-39}$ $\text{Fe}(\text{OH})_3 \sim 10^{-10}$
 B. Lead (II) Sulfide $K_{\text{sp}} = 3 \times 10^{-28}$ $\text{PbS} \sim 10^{-14}$
 C. Beryllium Hydroxide $K_{\text{sp}} = 7 \times 10^{-22}$ $\text{Be}(\text{OH})_2 \sim 10^{-7}$

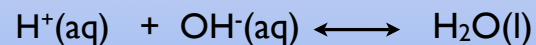
Neutralization

I can either have large concentrations of

H^+ or OH^-

but never both

The will reaction to get back to equilibrium



To solve we neutralize until all of one of them is gone

Acid no OH^-

Base no H^+

then we use the equilibrium expression to find the very small concentration left behind

You have a mixture of 100 mL of 1 M HCl
and 100 mL of 0.5 M NaOH

is this solution?

- A. Acidic ← more acid than base
look at moles!
- B. Basic
- C. Neutral
- HCl (1M)x(.1L) = .1 moles
NaOH (0.5M)x(.1L) = 0.05 moles

You have a mixture of 100 mL of 1 M HCl
and 100 mL of 0.5 M NaOH

How many moles of "excess" H^+ does this solution have?

- A. 0.1
- B. 0.05 ←
- C. 0.15
- D. 0
- HCl (1M)x(.1L) = .1 moles
NaOH (0.5M)x(.1L) = 0.05 moles
- .1 moles HCl = .1 moles H^+
.05 moles NaOH = .05 moles OH^-
they react until all the OH^- is gone
leaving .05 moles of H^+

You have a mixture of 100 mL of 1 M HCl
and 100 mL of 0.5 M NaOH

What is the $[H^+]$ of this solution

- A. 0.05/.1
- B. 0.05/.15
- C. 0.05/.2 ←
- D. 0.05/.05

.05 moles of H^+
volume is now 200 mL = 0.2 L
.05 moles/.2 L = .25 M

You have a mixture of 100 mL of 1 M HCl
and 100 mL of 0.5 M NaOH

What is the pH of this solution?

- A. -.6
- B. 0
- C. .6 ←
- D. 1

$[H^+] = 0.25 \text{ M}$
 $\text{pH} = -\log(.25) = -(-.6) = 0.6$

You have a mixture of 100 mL of 1 M HCl
and 100 mL of 0.5 M NaOH

What is the pOH of this solution?

- A. 12.4
- B. 13.4 ←
- C. 14.4
- D. -.6

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

You have a mixture of 100 mL of 1 M HCl
and 100 mL of 0.5 M NaOH

What is the $[OH^-]$?

- A. 10×13.4
- B. $10^{13.4}$
- C. $10^{-13.4}$ ←
- D. 1

$\text{pOH} = 13.4$
 $[OH^-] = 10^{-\text{pOH}}$