

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?

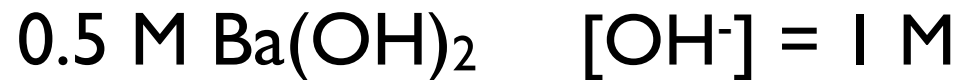
A. 0.5

B. 9

C. 14

D. 12

E. 10



[H⁺] = 10⁻¹⁴/1 = 10⁻¹⁴

pH = -log[10⁻¹⁴] = 14



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

A. 0 ←

B. 0.5 0.5 M Ba(OH)₂ [OH⁻] = 1 M

C. 1 pOH = -log[1] = 0

D. 12 pOH = 14 - pH

E. 14

Which of the following is the most acidic?

- A. acetic acid $K_a = 1.8 \times 10^{-5}$
- B. hydrofluoric acid $K_a = 7.2 \times 10^{-4}$ ← largest K_a
- 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 $K_a = 6.2 \times 10^{-10}$ ←

D. nitrite nitrous acid $K_a = 4.0 \times 10^{-4}$

smallest K_a
will be largest
 K_b

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}

D. 1.6×10^{-14}

E. 60×10^{-19}

← $K_b = K_w / K_a = 10^{-14} / 6 \times 10^{-5}$
 1.6×10^{-10}

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

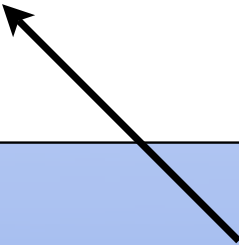
When you have only one compound in water
What you need to know if the “generic formula”

$$\text{MX} \quad K_{\text{sp}} = [\text{M}^+][\text{X}^-] = x^2 \quad x = (K_{\text{sp}})^{1/2}$$

$$\text{MX}_2 \quad K_{\text{sp}} = [\text{M}^+][\text{X}^-]^2 = 4x^3 \quad x = (K_{\text{sp}}/4)^{1/3}$$

$$\text{MX}_3 \quad K_{\text{sp}} = [\text{M}^+][\text{X}^-]^3 = 27x^4 \quad x = (K_{\text{sp}}/27)^{1/4}$$

Which is the most soluble?

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

Neutralization

I can either have large concentrations of



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

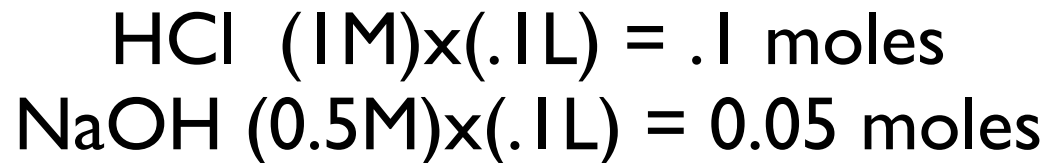
You have a mixture of 100 mL of 1 M HCl
and 100 mL of 0.5 M NaOH

is this solution?

A. Acidic ← more acid than base
look at moles!

B. Basic

C. Neutral



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 $\text{HCl } (1\text{M}) \times (.1\text{L}) = .1 \text{ moles}$
 $\text{NaOH } (0.5\text{M}) \times (.1\text{L}) = 0.05 \text{ moles}$
- B. 0.05 ←
- C. 0.15 $.1 \text{ moles HCl} = .1 \text{ moles H}^+$
 $.05 \text{ moles NaOH} = .05 \text{ moles OH}^-$
 they react until all the OH^- is gone
 leaving .05 moles of H^+
- D. 0

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 $[\text{OH}^-]$?

A. 10×13.4

B. $10^{13.4}$

C. $10^{-13.4}$ ←

D. 1

$$\begin{aligned} \text{pOH} &= 13.4 \\ [\text{OH}^-] &= 10^{-\text{pOH}} \end{aligned}$$