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

$0.5\ M\ Ba(OH)_2 \quad [OH^-] = 1\ M$

$[H^+] = 10^{-14}/1 = 10^{-14}$

$pH = -\log[10^{-14}] = 14$

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

A. 0

B. 0.5

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 $K_a = 7.2 \times 10^{-4}$ ← 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 Ka = 1.8×10^{-5}
B.	fluoride	hydrofluoric acid Ka = 7.2×10^{-4}
C.	cyanide	hydrocyanic acid Ka = 6.2×10^{-10} ←
D.	nitrite	nitrous acid Ka = 4.0×10^{-4}

smallest K_a will be largest K_b

Converting Ka to Kb

$$HA(aq) \longrightarrow H^+(aq) + A^-(aq)$$

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

$$A^-(aq) + H_2O(l) \longrightarrow HA(aq) + OH^-(aq)$$

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

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

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} $\longleftarrow$ $K_b = K_w / K_a = 10^{-14}/6\times10^{-5}$

1.6×10^{-10}

D. 1.6×10^{-14}

E. 60×10^{-19}

Which is the most soluble?

A. Aluminum Hydroxide $Al(OH)_3$ $K_{sp} = 3 \times 10^{-34}$

B. Barium Fluoride BaF_2 $K_{sp} = 1.8 \times 10^{-7}$ ←

C. Calcium Sulfate $Ca(SO_4)$ $K_{sp} = 5 \times 10^{-5}$

What is the solubility of ScF_3?

$$ScF_3\,(s) \longleftrightarrow Sc^{3+}(aq) + 3F^-(aq)$$

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

	ScF_3	Sc^{3+}	F^-
I	n_{solid}	0	0
C	-x	+x	+3x
E	n-x	+x	+3x

$K = [Sc^{3+}][F^-]^3$

$K = (x)(3x)^3$

$K = 27x^4 = 4.2 \times 10^{-18}$

$x = 1.99 \times 10^{-5}$

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

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

Which is the most soluble?

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

Neutralization

I can either have large concentrations of

H^+ or OH^-

but never both

The will reaction to get back to equilibrium

$$H^+(aq) + OH^-(aq) \longleftrightarrow H_2O(l)$$

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 ←

B. Basic

C. Neutral

more acid than base
look at moles!

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

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

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

$$pOH = 14 - pH$$

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 x 13.4

B. $10^{13.4}$

C. $10^{-13.4}$ ←

D. 1

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