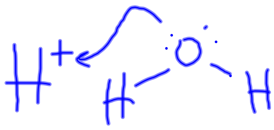
#11-16

12.

b) $pH = -\log[H^+_{(aq)}]$

$pH = -\log[H_3O^+_{(aq)}]$



$pH = -\log[0.045]$

$\boxed{pH = 1.35}$

15. a) $[OH^-_{(aq)}] = 4.3 \times 10^{-5} M$

$pH = ?$

$K_w = [H^+_{(aq)}][OH^-_{(aq)}]$

$[H^+_{(aq)}] = \dfrac{K_w}{[OH^-_{(aq)}]}$

$[H^+_{(aq)}] = \dfrac{1.0 \times 10^{-14}}{4.3 \times 10^{-5} M}$

$[H^+_{(aq)}] = 2.3 \times 10^{-10} M$

$pOH = -\log[OH^-_{(aq)}]$

$pOH = -\log[4.3 \times 10^{-5}]$

$\underline{pOH = 4.37}$

$pH + pOH = 14.00$

$pH = 14.00 - 4.37$

$pH = 9.63$

$pH = -\log[H^+_{(aq)}]$

$pH = -\log[2.3 \times 10^{-10}]$

$\boxed{pH = 9.64}$

$$K_w = [H^+_{(aq)}][OH^-_{(aq)}] = 1.0 \times 10^{-14}$$

$$pH = -\log[H^+_{(aq)}]$$

$$[H^+_{(aq)}] = 10^{-pH}$$

$$pOH = -\log[OH^-_{(aq)}]$$

$$[OH^-_{(aq)}] = 10^{-pOH}$$

$$pH + pOH = 14.00$$

Strong Acids

Calculate the concentration of the hydroxide ions, pH and pOH of a 0.15 mol/L solution of sulfuric acid at 25°C.

Strong acids will always completely ionize

$$H_2SO_{4(aq)} \longrightarrow 2H^+_{(aq)} + SO^{2-}_{4(aq)}$$

0.15mol/L $\boxed{0.30\text{mol/L}}$

$[OH^-_{(aq)}] = ?$

$pH = ?$

$pOH = ?$

$pH = -\log[H^+_{(aq)}]$

$pH = -\log[0.30]$

$\boxed{pH = 0.52}$

$pH + pOH = 14.00$

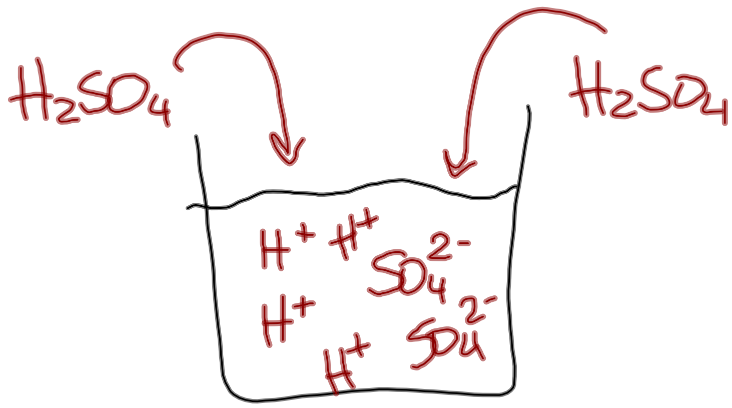
$pOH = 14.00 - 0.52$

$\boxed{pOH = 13.48}$

$[OH^-_{(aq)}] = 10^{-pOH}$

$[OH^-_{(aq)}] = 10^{-13.48}$

$\boxed{[OH^-_{(aq)}] = 3.3 \times 10^{-14} M}$



H_2SO_4
H_2SO_4
H^+ H^+ SO_4^{2-}
H^+
H^+ SO_4^{2-}

$$\% \, rxn = \frac{exp}{theor} \times 100\%$$

$$exp. = theor.$$

Strong Bases (Ionic Hydroxides)

Calculate the hydrogen ion concentration, pH and pOH of a 0.25 mol/L solution of barium hydroxide.

$$Ba(OH)_{2(s)} \longrightarrow Ba^{2+}_{(aq)} + 2OH^-_{(aq)}$$

0.25 mol/L   0.50 mol/L

$[H^+_{(aq)}] = ?$
$pH = ?$
$pOH = ?$

$pOH = -\log[OH^-_{(aq)}]$
$pOH = -\log[0.50]$
$\boxed{pOH = 0.30}$

$pH + pOH = 14.00$
$pH = 14.00 - 0.30$
$\boxed{pH = 13.70}$

$[H^+_{(aq)}] = 10^{-pH}$
$[H^+_{(aq)}] = 10^{-13.70}$
$\boxed{[H^+_{(aq)}] = 2.0 \times 10^{-14} M}$

Water Equilibrium Worksheet

Ionization Constants for Acids

Strong acids - ionizes quantitatively in water to form hydronium ions

$$HCl_{(aq)} + H_2O_{(l)} \longrightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

Weak acids - ionizes partially in water to form hydronium ions

$$CH_3COOH_{(aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$$

To describe the equilibrium of acids in water, the equilibrium law is used to calculate the acid ionization constant, K_a.

$$CH_3COOH_{(aq)} + H_2O_{(l)} \longleftrightarrow H_3O^+_{(aq)} + CH_3COO^-_{(aq)}$$

$$K_a = 1.8 \times 10^{-5} \text{ mol/L}$$

Ex. Predict the hydrogen ion concentration, and pH of a 1.0 mol/L acetic acid solution at equilibrium.

Ex. The pH of a 0.25 mol/L carbonic acid solution at equilibrium is found to be 3.48. Calculate the K_a.

Weak Bases

Weak bases react with water to form the hydroxide ion and conjugate acid of the base.

$$NH_{3(aq)} + H_2O_{(l)}$$

***Eqm greatly favours reverse reaction**

$$K_{eq} =$$

$$K_b =$$

base dissociation constant