	Acids		Bases
2. a)	$HClO_4$	$\rightarrow$	ClO_4^-
	NH_4^+	$\rightarrow$	NH_3
	H_2O	$\rightarrow$	OH^-
	HCO_3^-	$\rightarrow$	CO_3^{2-}
3.	$-CN^-$	$\rightarrow$	HCN
	SO_4^{2-}	$\rightarrow$	HSO_4^-
	H_2O	$\rightarrow$	H_3O^+
	HCO_3^-	$\rightarrow$	H_2CO_3

4.

$$\underset{A}{\boxed{HSO_{3(aq)}^-}} + \underset{B}{\boxed{H_2O_{(l)}}} \rightleftharpoons \boxed{SO_{3(aq)}^{2-}} + \boxed{H_3O_{(aq)}^+}$$

H^+ H^+

$$K_a = \frac{[SO_{3(aq)}^{2-}][H_3O_{(aq)}^+]}{[HSO_{3(aq)}^-][H_2O_{(l)}]} = 1.23 \times 10^{-7}$$

Water Equilibrium

Conductivity is due to the presence of ions. For water:

$$H_2O_{(l)} \Longleftrightarrow H^+_{(aq)} + OH^-_{(aq)}$$

- therefore $K = \dfrac{[H^+][OH^-]}{[H_2O]}$ is very small

- slight conductivity shows that equilibrium greatly favors water molecules (less than 2 H^+ per billion water)

- therefore the concentration of water in pure water and in dilute aqueous solutions is essentially constant and can be combined with the equilibrium constant to produce a new constant called the *Ion Product Constant*

Ionization Constant for water (ion product constant)

$\mathbf{K_w = [H^+]\,[OH^-] = 1.0 \times 10^{-14}}$ at SATP

Since $[H^+]$ and $[OH^-]$ are found in 1:1 ratio
($H_2O_{(l)} \Longleftrightarrow H^+_{(aq)} + OH^-_{(aq)}$)

$[H^+_{(aq)}] = [OH^-_{(aq)}] = 1.0 \times 10^{-7}$ mol/L in **neutral** solutions.

Arrhenius's Theory - acid is a substance that ionizes water t produce H^+ ions.
- additional ions produced by the acid increases the H^+ concentration in the water. (more acid, more H^+)

Therefore acids always have a $[H^+] > 10^{-7}$ mol/L

Basic solutions produce a $[OH^-]$ greater than 10^{-7} mol/L

K_w can be used to calculate either $[H^+]$ or $[OH^-]$

since $\mathbf{K_w = [H^+]\,[OH^-]}$ then $[H^+] = K_w / [OH^-]$

and $[OH^-] = K_w / [H^+]$

	H^+
Acid	x
water	1.0×10^{-7} mol/L
	$> 1.0 \times 10^{-7}$ mol/L

pH and pOH

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

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

$pH + pOH = 14.00$ $K_w = [H^+_{(aq)}][OH^-_{(aq)}] = 1.0 \times 10^{-14}$

Ex. Calculate the pH of a solution where $[H^+_{(aq)}] = 3.24 \times 10^{-4}$ M.

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

$pH = -\log[3.24 \times 10^{-4}]$

$pH = 3.489$

Ex. Calculate the concentration of hydroxide ions in a solutic with a pOH of 10.14.

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

$\boxed{10}\boxed{\wedge}($

$[OH^-_{(aq)}] = 10^{-10.14}$

$\boxed{[OH^-_{(aq)}] = 7.2 \times 10^{-11} M}$

Strong Acids

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

Strong acids will always completely ionize

$$HCl_{(aq)} \longrightarrow H^+_{(aq)} + Cl^-_{(aq)}$$

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

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

$pH = -\log[0.15]$

$\boxed{pH = 0.82}$

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

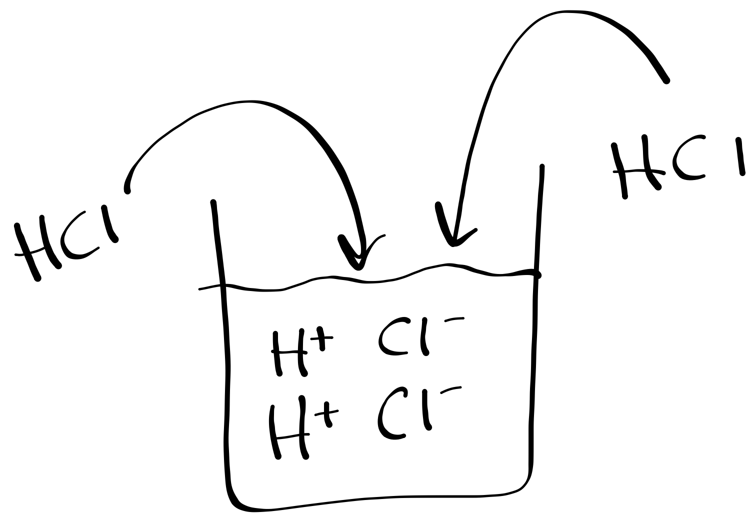
$[OH^-_{(aq)}] = \dfrac{1.0 \times 10^{-14}}{0.15}$

$\boxed{[OH^-_{(aq)}] = 6.7 \times 10^{-14}\text{ M}}$

$pOH = -\log[OH^-]$

$pOH = -\log[6.7 \times 10^{-14}]$

$\boxed{pOH = 13.17}$


HCl
HCl
H^+ Cl^-
H^+ Cl^-

Strong Bases
(Ionic Hydroxides)

Calculate the hydrogen ion concentration in a 0.25 mol/L solution of barium hydroxide.

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

0.25 mol/L $\qquad$ $\boxed{0.50\ mol/L}$

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

$$[H^+_{(aq)}] = \frac{K_w}{[OH^-_{(aq)}]} = \frac{1.0 \times 10^{-14}}{0.50\ M} = \boxed{2.0 \times 10^{-14}\ M}$$

H^+ donors

$$\overset{A}{\boxed{HF_{(aq)}}} + \overset{B}{\boxed{H_2O_{(l)}}} \rightleftharpoons \overset{B}{\boxed{F^-_{(aq)}}} + \overset{A}{\boxed{H_3O^+_{(aq)}}}$$

H^+ H^+