With seawater, it contains acid H_2CO_3, which dissociates

① $H_2CO_3 \rightleftharpoons H^+ + HCO_3^-$

② $HCO_3^- \rightleftharpoons H^+ + CO_3^{2-}$

③ $HCO_3^- + OH^- \rightleftharpoons H_2O + CO_3^{2-}$

Notice that adding H^+ and OH^-, adding something acidic/basic, will shift to making HCO_3^- rather than add more H^+/OH^-. Therefore pH doesn't change very much

Buffers: a substance that helps a solution resist pH change

- contains almost/relatively high and equal concentrations of both the acid and its conjugate base

↳ adding a small amount of acid/base only slightly neutralizes one component of the buffer = pH doesn't change much

$$HA \rightleftharpoons H^+ + A^-$$

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

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

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

$$-\log[H^+] = -\log K_a + \log \frac{[A^-]}{[HA]} \quad \text{(log laws)}$$

$-\log[H^+]$ = pH, $-\log K_a$ = pKa

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

= Henderson-Hasselbalch Equation (for buffer solns @ equilibrium)

- This equation is ONLY for buffers

Calc. pH of soln containing 0.22 M acetic acid and 0.13 M sodium acetate (K_a acetic acid = 1.8×10^{-5})

	$HC_2H_3O_2$	$\rightleftharpoons$	H^+	+	$C_2H_3O_2^-$
I	0.22		0		0.13
C	−x		+x		+x
E	0.22 − x		x		0.13 + x
	≈ 0.22				≈ 0.13

↳ since we take the "ignore x approx." when calculating x, we are saying the equilibrium is pretty much equal to its initial concentration because x is so small

HH Equation method:

$$pH = pK_a + \log \frac{[A^-]}{[HA]}$$

$$pH = (-\log(1.8 \times 10^{-5}) + \log \frac{(0.13)}{(0.22)}$$

$$pH = 4.74 + \log \frac{0.13}{0.22} = \boxed{4.51}$$

If $[A^-] = [HA]$, then

$$pH = pK_a + \log \frac{[A^-]}{[HA]} = pK_a + \log(1) = pK_a$$

Under what conditions makes a buffer good?

- $HA \rightleftharpoons H^+ + A^-$
  - ideal pH of a buffer is when $pH = pK_a$
    - ↳ when $[HA] = [A^-]$ $\quad \frac{[A^-]}{[HA]}$
- Buffer is effective buffer when ratio is in the range of 0.1 to 10
  - ↳ pH range ± 1 from pK_a

- Capacity of Buffer = how much acid can you add and still have an effective buffer?
  - ↳ the greater the initial concentrations of HA and A^-, the stronger the acid/base can be added without changing the pH

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

1 M A^-	0.1 M A^-	0.01 M A^-
1 M HA	0.1 M HA	0.01 M HA

↳ largest capacity

- ↳ if H^+ is more than buffer, pH will change more drastically

- Buffer reacts with H^+ to avoid pH change

find # moles of sodium acetate that must be added to 250.0 mL of 0.16 M acetic acid in order to prepare a pH 4.70 buffer ($K_a = 1.76 \times 10^{-5}$)

$$pH = pK_a + \log \frac{[A^-]}{[HA]} \qquad K_a = 1.76 \times 10^{-5} \qquad pK_a = 4.75$$

$$4.70 = 4.75 + \log \frac{[A^-]}{0.16}$$

$$-0.05 = \log \frac{[A^-]}{0.16}$$

$$10^{-0.05} = \frac{[A^-]}{0.16}$$

$$(0.16)(10^{-0.05}) = [A^-] = 0.143\ M$$

$(0.143\ M)(0.250\ L) = 0.0358\ mol$ sodium acetate (CH_3COO^-)