- **Blood Buffering**
  - Our blood-buffering system is much more effective at buffering against added H^+ than added OH^- because there is more of the conjugate base (bicarbonate: carbonic acid = 20:1)

- **Picking a Buffer**
    - **Q1: You need to make a buffer solution at pH = 7.40. In order to make the best buffer possible (i.e., effectively buffers against both H^+ and OH^-), which would be the best choice?**
      When $[A^-] = [HA]$, the buffer is optimal.
      $pH = pK_a + \log \frac{[A^-]}{[HA]}$
      $pH = pK_a + \log (1)$
      So $pH = pK_a$
    - When challenge with 0.5 mL of 1M HCl,
      1M, 0.1M, 0.01M Tris all had the same initial pH
      ($[HA] / [A^-]$ was 1:1 for all pH=8)
      - Initial pH is a result of $[A^-]/[HA]$ ratio
      - Buffer capacity decreased with concentrations of A^- and HA
    - A good buffer has a pK_a close to the desired pH.
    - Excess acid can be exhaled as CO_2; excess base excreted as bicarb.
    - Rule of thumb: Choose a buffer with a pK_a within 1 pH unit of the desired pH.
      - Buffer at pH = 9.0, choose boric acid, pKa = 9.24
        1. Need conjugate base, borate (sodium borate)
        2. Add NaOH (strong base) to boric acid to make borate
           $pH = pK_a + \log \frac{[A^-]}{[HA]}$
           $9.0 = 9.24 + \log \frac{[A^-]}{[HA]}$
           Solve for $[A^-] : [HA]$

- **Henderson–Hasselbalch equation**
  - **Calculating pH**
    - What is the pH of a buffer that is 0.12 *M* in lactic acid and 0.10 *M* in sodium lactate? K_a for lactic acid is 1.4×10^{-4}.
      $pH = pK_a + \log \frac{[A^-]}{[HA]}$
      $= -\log (1.4 \times 10^{-4}) + \log \frac{0.10\ M}{0.12\ M}$
      $= 3.85 + (-0.08) = 3.77$
  - **Calculating ratio of $[A^-]/[HA]$**
    - To make a buffer solution at pH = 7.40, what ratio of $[A^-]/[HA]$ do you need in your phosphate buffer? $pK_a = 7.21$.
      $pH = pK_a + \log \frac{[A^-]}{[HA]}$
      $7.40 = 7.21 + \log \frac{[A^-]}{}$

$$0.19 = \log \frac{[A^-]}{[HA]} \quad \text{so } \frac{[A^-]}{[HA]} = 1.5$$

- **Calculating pK_a**
  - A solution of 0.225 M $NaNO_2$ and 1.0 M HNO_2 has a pH = 2.51. What is the pK_a of HNO_2?

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

    $$2.51 = pK_a + \log \frac{(0.225)}{(1.0)}$$

    $$2.51 = pK_a - 0.65 \qquad \text{so } pK_a = 3.16$$

- **Adding a Strong Acid or Base to a Buffer**
  - Adding a strong acid/base converts the weak base/acid to its conjugate acid/base.
  - Find [HA] and [A⁻] when the added acid/base reacts.
  - Use the Henderson–Hasselbalch equation to find pH.