1. Calculate the pH of a $1.0 \times 10^{-3}$ M solution of hydrochloric acid (HCl)

   HCl is a strong acid: \[ \text{HCl} \rightarrow \text{H}^+ + \text{Cl}^- \]

   \[ [\text{H}^+] = 1.0 \times 10^{-3} \rightarrow \text{pH} = -\log([\text{H}^+]) = -\log(10^{-3}) = \boxed{3} \]

2. Calculate the pH of a 0.9 M solution of hydrobromic acid (HBr)

   HBr is a strong acid: \[ \text{HBr} \rightarrow \text{H}^+ + \text{Br}^- \]

   \[ [\text{H}^+] = 0.9 \times 10^{-1} \text{ M} \rightarrow \text{pH} = -\log(0.9) = \boxed{0.046} \]
   (very acidic!)

3. Calculate the pH of a $2.234 \times 10^{-6}$ M solution of potassium hydroxide (KOH).

   KOH is a strong base: \[ \text{KOH} \rightarrow \text{K}^+ + \text{OH}^- \]

   \[ [\text{OH}^-] = 2.234 \times 10^{-6} \text{ M} \rightarrow \text{pOH} = 5.65 \]

   \[ \text{pH} = 14 - \text{pOH} = \boxed{8.35} \]

4. Calculate the pH of 735 liters of a solution containing 0.34 moles of nitric acid (HNO₃).

   HNO₃ is a strong acid: \[ \text{HNO}_3 \rightarrow \text{H}^+ + \text{NO}_3^- \]

   First we need the concentration:

   \[ \frac{0.34 \text{ moles HNO}_3}{0.735 \text{ L}} = 0.4626 \text{ M} \]

   So \[ [\text{H}^+] = 0.4626 \text{ M} \rightarrow \text{pH} = 0.33 \]
   (very acidic)
5. Calculate the pH of a 2.0 liter solution containing 0.005 g of HCl.

$$\text{concentration: } \frac{0.005 \text{ g}}{37 \text{ g HCl}} \left( \frac{1 \text{ mol}}{2 \text{ L}} \right) = \frac{0.005}{37(2)} \text{ M} = 0.76 \times 10^{-5} \text{ M}$$

$$[H^+] = 0.76 \times 10^{-5} \text{ M (strong acid)} \Rightarrow \text{pH} = 4.17$$

6. Calculate the pH and pOH of a solution with a volume of 5.4 liters that contains 15 g of hydrochloric acid (HCl) and 25 g of nitric acid (HNO₃).

Assume that both strong acids dissociate completely.

$$\text{HCl} \rightarrow H^+ + Cl^- \quad 15 \text{ g HCl} \left( \frac{1 \text{ mol}}{37 \text{ g HCl}} \right) = 0.405 \text{ moles H}^+$$

$$\text{HNO}_3 \rightarrow H^+ + NO_3^- \quad 25 \text{ g HNO}_3 \left( \frac{1 \text{ mol}}{48 \text{ g HNO}_3} \right) = 0.397 \text{ moles H}^+$$

$$[H^+] = \frac{(0.405 + 0.397) \text{ mol}}{5.4 \text{ L}} = 0.148 \text{ M} \Rightarrow \text{pH} = 0.83$$

7. A swimming pool has a volume of 1 million liters. How many grams of HCl would you need to add to the pool to bring the pH from 7 down to 4? Assume that the volume of the HCl is negligible.

$$\text{pH} = -\log [H^+] \Rightarrow \text{at pH} = 7, \quad [H^+] = 10^{-7} \text{ M} \quad \text{at pH} = 4, \quad [H^+] = 10^{-4} \text{ M}$$

$$\frac{10^{-7} \text{ mol}}{L} \times (10^6 \text{ L}) = 0.1 \text{ mol} \quad \text{of HCl (strong acid)} \quad \text{difference} \approx 100 \text{ moles}.$$  

$$100 \text{ moles HCl} \left( \frac{37 \text{ g HCl}}{1 \text{ mol}} \right) = 3700 \text{ g} = 3.7 \text{ kg of HCl}$$

8. Calculate the pH and pOH of a solution that was made by adding 400 ml of water to 350 ml of a 5.0 x 10⁻³ M NaOH solution.

$$5.0 \times 10^{-3} \text{ moles} (0.35 \text{ mol}) = 0.00175 \text{ moles NaOH} \rightarrow \frac{0.00175 \text{ mol}}{(0.35 + 0.4) \text{ L}}$$

$$\text{pOH} = -\log (0.00233) = 2.63$$

$$\text{pH} = 14 - \text{pOH} = 11.37$$
9. The active ingredient of aspirin is acetyl salicylic acid (C₈H₇O₂COOH), which has $K_a = 3.0 \times 10^{-4}$. Acetyl salicylic acid is often abbreviated "ASA".

(a) Calculate the pH of a solution made by dissolving 500 mg of ASA in water, and diluting it to 50 ml.
(b) Repeat the calculation, this time dissolving the ASA in 1000 ml of water.
(c) Is the pH higher or lower in the more dilute solution? In which solution is the fraction of ionized ASA higher?

Let's abbreviate it: $\text{HA} \rightleftharpoons \text{H}^+ + \text{A}^-$.

\[
K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} = 3.0 \times 10^{-4}
\]

\[\begin{align*}
\text{500 mg} & = 0.5 \text{ g} ; \text{ MW} = 9(12) + 8(16) + 8(1) = 180 \text{ g/mol}, \\
0.5 \text{ g} \left( \frac{1 \text{ mol}}{180 \text{ g}} \right) & = 0.0027778 \text{ mol} \text{ ASA} \\
& \rightarrow 0.0027778 \text{ mol} \text{ ASA} \\
& = 0.556 \text{ M ASA}
\end{align*}\]

Now the pH:

\[\frac{0.556 - x}{x} \rightarrow \frac{3 \times 10^{-4}}{0.556} \rightarrow x = 0.0041 \text{ M H}^+ \]

\[\Rightarrow \text{pH} = 2.4\]

\[\begin{align*}
\text{(b)} & \quad \frac{0.0027778 \text{ mol}}{1 \text{ L}} = 0.0027778 \text{ M H}^+ \rightarrow 3 \times 10^{-4} = \frac{x^2}{0.0027778} \\
x & \approx 9.13 \times 10^{-4} \text{ M H}^+ \\
\Rightarrow \text{pH} & \approx 3.04
\end{align*}\]

(c) The fraction of ionized ASA is higher in the more conc. soln.
10. Papaverine hydrochloride (pap-HCl) is a salt of a weak base (papaverine) and a strong acid (HCl). It is a drug used as a muscle relaxant. Pap-HCl is a weak acid overall. At 25°C, a 205 mM solution of pap-HCl has a pH of 3.31. Compute the $K_a$ of pap-HCl.

Let's abbreviate the weak acid as "HA"; $HA \rightleftharpoons H^+ + A^-$

\[
205 \text{ mM} = 0.205 \text{ M} \quad \text{pH} = 3.31 \implies [H^+] = 10^{-3.31} = 4.89 \times 10^{-4} \text{ M}
\]

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\text{now: } K_a = \frac{[H^+][A^-]}{[HA]} = \frac{(4.89 \times 10^{-4})(4.89 \times 10^{-4})}{0.205 - 4.89 \times 10^{-4}}
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\[
K_a = 1.17 \times 10^{-6}
\]

Note: $K_a = 1.16 \times 10^{-6}$ if we ignore this.

11. A 0.040 M solution of a weak acid (call it HA) has a pH of 4.70. Calculate the $K_a$ and $pK_a$ of the acid.

\[
HA \rightleftharpoons H^+ + A^-
\]

\[
K_a = \frac{[H^+][A^-]}{[HA]} \approx \frac{10^{-4.7}}{0.040} = 2.5 \times 10^{-5}
\]

\[
\rho K_a = -\log K_a = -\log (2.5 \times 10^{-5}) = 7.8
\]

\[
K_a = 2.5 \times 10^{-5}; \quad \rho K_a = 7.8
\]