Dilution Calculations: Water, water everywhere

Dilution: adding water to a concentrated solution to decrease the molarity

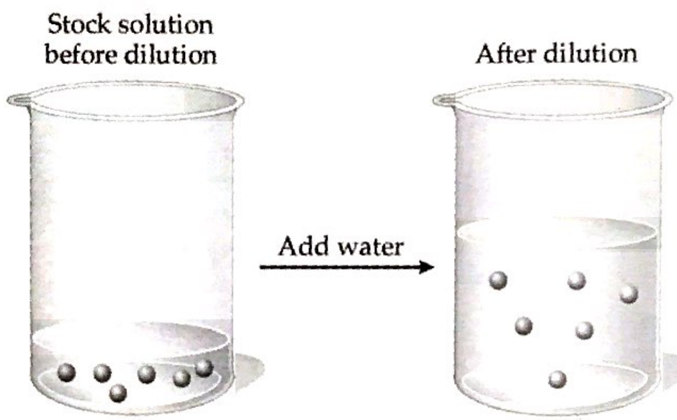
- **Stock solutions:** concentrated solutions that labs purchase to save time and space in the laboratory.
- H_2O is then added to the stock solution to achieve the desired molarity.

When diluting solutions with pure H_2O it is important to remember the number of moles does not change.

The Dilution Equation

$$\underbrace{M_1V_1 = M_2V_2}$$

Can use mL or L for volume!



The two beakers contain the same number of moles of solute.

For the dilution of acidic and basic solutions, it's important to be able to do simple estimations without a calculator!

For example,

1. Doubling the volume of solution by adding an amount of distilled water equal to the original volume of solution will decrease the concentration by half. $(\times \frac{1}{2})$
2. Increasing the volume of solution by adding 9 parts distilled water to 1 part original solution with decrease the concentration by a factor of 10. $(\times \frac{1}{10})$

Guided Practice

1. Consider a 200. mL solution of 0.200 M HBr after the addition of 200. mL of distilled water.
   a. What will the new concentration of HBr be?

$$(0.200M)(200.mL) = x(400.mL) \Rightarrow x = \frac{(0.200)(200)}{400.} = \frac{0.200}{2} = \boxed{0.100M = [HBr]}$$

   b. What will be the pH of the new solution?

$$pH = -\log(1.00E-1) = \boxed{1.000} \text{ (strong acid, so } [HBr] = [H_3O^+])$$

2. Consider a 100. mL solution of 1.0 M LiOH after the addition of 900. mL of distilled water.
   a. What will the new concentration of LiOH be?

$$(1.0M)(100.mL) = x(1,000.mL) \Rightarrow x = \frac{(1.0)(100.)}{(1,000.)} = \frac{1.0}{10} = \boxed{0.10M = [LiOH]}$$

   b. What will be the pH of the new solution?

$$pOH = -\log[OH^-] = -\log(1.0E-1) = 1.00 \text{ (strong base, so } [LiOH] = [OH^-])$$

$$pH = 14 - 1.00 = \boxed{13.00}$$