

Solutions and Concentrations

Example 1: If you dissolve 25.5g KBr in enough water to make 1.75 L of solution, what is the molarity of the solution?

$$M_w(\text{KBr}) = 39.10 \text{ g/mol} + 79.90 \text{ g/mol} = 119.0 \text{ g/mol}$$

$$\# \text{ moles (KBr)} = 25.5 \text{ g} \cdot \frac{1 \text{ mole}}{119.0 \text{ g}} = 0.214 \text{ mole}$$

$$\text{molarity (M)} = \frac{0.214 \text{ mole}}{1.75 \text{ L}} = 0.122 \text{ M}$$

Example 2: How many liters of a 0.125 M NaOH solution contain 0.255 mole NaOH?

$$M = \frac{n}{V} \Rightarrow V = \frac{n}{M}$$

$$V = \frac{0.255 \text{ mole}}{0.125 \text{ mol/L}} = 2.04 \text{ L}$$

Example 3: What is the molarity of a 150 ppm solution of NaCl ($d = 1 \text{ g/cm}^3$)

$$150 \text{ ppm} = \frac{150 \text{ mg (NaCl)}}{1 \text{ kg (water)}} \rightarrow \begin{array}{l} \text{convert to} \\ \text{moles} \end{array} \rightarrow \text{volume}$$

$$M_w(\text{NaCl}) = 58.44 \text{ g/mole}$$

$$\begin{aligned} \# \text{ moles (NaCl)} &= 150 \text{ mg} \cdot \frac{1 \text{ g}}{1000 \text{ mg}} \cdot \frac{1 \text{ mole}}{58.44 \text{ g}} \\ &= 2.92 \cdot 10^{-3} \text{ mole} \end{aligned}$$

$$\text{mass (solution)} = 1000.15 \text{ g} \approx 1000 \text{ g} \quad \begin{array}{l} 150 \text{ mg} \cdot \frac{1 \text{ g}}{1000 \text{ mg}} = \\ 0.15 \text{ g} \end{array}$$

$$1 \text{ g/cm}^3 = 1 \frac{\text{g}}{\text{cm}^3} \cdot \frac{1 \text{ mL}}{1 \text{ cm}^3} = \boxed{1 \text{ g/mL}}$$

$$\begin{aligned} 1 \text{ g/cm}^3 &= 1 \frac{\text{g}}{\text{cm}^3} \cdot \frac{1000 \text{ cm}^3}{1 \text{ dm}^3} = \frac{1000 \text{ g}}{1 \text{ dm}^3} \cdot \frac{1 \text{ dm}^3}{1 \text{ L}} \\ &= \frac{1 \text{ kg}}{1 \text{ L}} \end{aligned}$$

$$\begin{aligned} \text{molarity (NaCl)} &= \frac{2.92 \cdot 10^{-3} \text{ mole}}{1 \text{ L}} = 2.92 \cdot 10^{-3} \text{ M} \\ &= 2.92 \cdot 10^{-3} \text{ M} \cdot \frac{1000 \text{ mM}}{1 \text{ M}} \\ &= 2.92 \text{ mM} \end{aligned}$$

What is the molality of a solution prepared by dissolving 32.0 g CaCl_2 in 271 g of H_2O ?

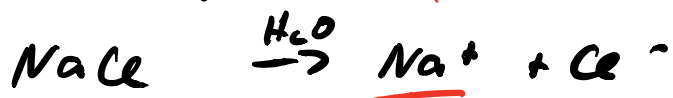
$$\# \text{ moles } (\text{CaCl}_2) = 32.0 \text{ g} \cdot \frac{1 \text{ mole}}{110.98 \text{ g}} = 0.288 \text{ mole}$$

$$m = \frac{0.288 \text{ mole } (\text{CaCl}_2)}{271 \text{ g} \cdot \frac{1 \text{ kg}}{1000 \text{ g}}} = 1.06 \text{ m}$$

Common Ion Problems

What is the concentration of Na^+ when you mix 100 mL of 0.100 M Na_2SO_4 and 0.250 L of 0.500 M NaCl solution (both are strong electrolytes)

1. Identify common ion



Na^+ is the common ion

2. Calculate # of moles of Na^+ in both solutions

Na_2SO_4 :

$$\begin{aligned}\# \text{ moles } (\text{Na}^+) &= 2 \cdot 0.100 \text{ M} \cdot 100 \text{ mL} \cdot \frac{1 \text{ L}}{1000 \text{ mL}} \\ &= 0.0200 \text{ mol}\end{aligned}$$

NaCl

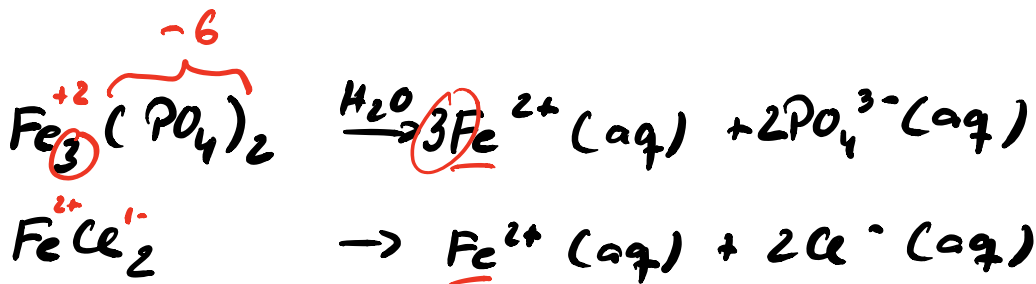
$$\# \text{ moles } (\text{Na}^+) = 1 \cdot 0.500 \text{ M} \cdot 0.250 \text{ L} = 0.125 \text{ mol}$$

$$\begin{aligned}\text{Total moles } (\text{Na}^+) &= 0.0200 \text{ mol} + 0.125 \text{ mol} \\ &= 0.145 \text{ mol}\end{aligned}$$

3. Calculate the molarity of Na^+

$$\text{Total Volume} = 100 \text{ mL} \cdot \frac{1 \text{ L}}{1000 \text{ mL}} + 0.250 \text{ L} = 0.350 \text{ L}$$

$$\text{M } (\text{Na}^+) = \frac{0.145 \text{ mol}}{0.350 \text{ L}} = 0.414 \text{ M}$$



Dilution: To what volume should you dilute 0.200 L of a 15.0 M NaOH solution to obtain a 3.00 M NaOH solution?

Given: $V_i = 0.200 \text{ L}$
 $M_i = 15.0 \text{ M}$
 $M_f = 3.00 \text{ M}$

$$V_i M_i = V_f M_f$$

$$V_f = \frac{V_i M_i}{M_f} = \frac{0.200 \text{ L} \cdot 15.0 \text{ M}}{3.00 \text{ M}} = 1 \text{ L}$$