

## REACTIONS IN SOLUTION

Section 5.5



## Terminology

In solution we need to define the -

- **SOLVENT**

the component whose physical state is preserved when solution forms

- **SOLUTE**

the other solution component



## Concentration of Solute

The amount of solute in a solution is given by its **concentration**.

$$\text{Molarity (M)} = \frac{\text{moles solute}}{\text{liters of solution}}$$



**PROBLEM:** Dissolve 5.00 g of  $\text{NiCl}_2 \cdot 6\text{H}_2\text{O}$  in enough water to make 250 mL of solution. Calculate molarity.

**Step 1:** Calculate moles of  $\text{NiCl}_2 \cdot 6\text{H}_2\text{O}$

$$5.00 \text{ g} \cdot \frac{1 \text{ mol}}{237.7 \text{ g}} = 0.0210 \text{ mol}$$

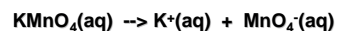
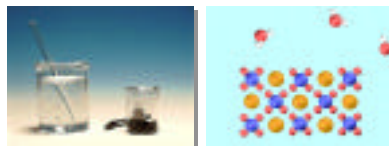
**Step 2:** Calculate molarity

$$\frac{0.0210 \text{ mol}}{0.250 \text{ L}} = 0.0841 \text{ M}$$

$$[\text{NiCl}_2 \cdot 6\text{H}_2\text{O}] = 0.0841 \text{ M}$$



## The Nature of the $\text{KMnO}_4$ Solution



If  $[\text{KMnO}_4] = 0.30 \text{ M}$ , then

$$[\text{K}^+] = [\text{MnO}_4^-] = 0.30 \text{ M}$$

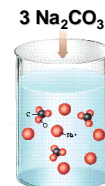
## The Nature of a $\text{Na}_2\text{CO}_3$ Solution

This water-soluble compound is ionic  
 $\text{Na}_2\text{CO}_3(\text{aq}) \rightarrow 2 \text{Na}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$

If  $[\text{Na}_2\text{CO}_3] = 0.100 \text{ M}$ , then

$$[\text{Na}^+] = 0.200 \text{ M}$$

$$[\text{CO}_3^{2-}] = 0.100 \text{ M}$$



## USING MOLARITY

What mass of oxalic acid,  $\text{H}_2\text{C}_2\text{O}_4$ , is required to make 250. mL of a 0.0500 M solution?

Because  
Conc (M) = moles/volume = mol/V  
this means that

$$\text{moles} = M \cdot V$$



## USING MOLARITY

What mass of oxalic acid,  $\text{H}_2\text{C}_2\text{O}_4$ , is required to make 250. mL of a 0.0500 M solution?

$$\text{moles} = M \cdot V$$

**Step 1:** Calculate moles of acid required.

$$(0.0500 \text{ mol/L})(0.250 \text{ L}) = 0.0125 \text{ mol}$$

**Step 2:** Calculate mass of acid required.

$$(0.0125 \text{ mol})(90.00 \text{ g/mol}) = 1.13 \text{ g}$$

## Preparing Solutions



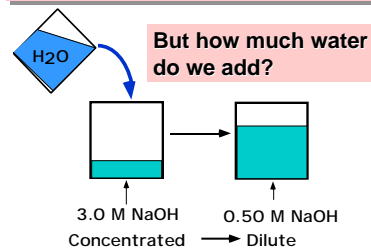
- Weigh out a solid solute and dissolve in a given quantity of solvent.
- Dilute a concentrated solution to give one that is less concentrated.

**PROBLEM:** You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What do you do?

Add water to the 3.0 M solution to lower its concentration to 0.50 M

**Dilute the solution!**

**PROBLEM:** You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What do you do?



**PROBLEM:** You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What do you do?

How much water is added?

The important point is that --->

moles of NaOH in ORIGINAL solution =  
moles of NaOH in FINAL solution

**PROBLEM:** You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What do you do?

Moles of NaOH in original solution =

$$M \cdot V =$$

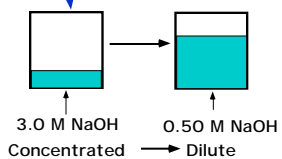
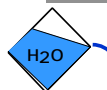
$$(3.0 \text{ mol/L})(0.050 \text{ L}) = 0.15 \text{ mol NaOH}$$

Therefore, moles of NaOH in final solution must also = 0.15 mol NaOH

$$(0.15 \text{ mol NaOH})(1 \text{ L}/0.50 \text{ mol}) = 0.30 \text{ L}$$

or **300 mL = volume of final solution**

**PROBLEM:** You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What do you do?



**Conclusion:**

**add 250 mL of water to 50.0 mL of 3.0 M NaOH to make 300 mL of 0.50 M NaOH.**

## Preparing Solutions by Dilution



A shortcut

$$M_{\text{initial}} \cdot V_{\text{initial}} = M_{\text{final}} \cdot V_{\text{final}}$$

