Stoichiometry of Solutions

9.3

Learning Goals

By the end of this lesson students will be able to:

- Use stoichiometry to solve problems involving solutions and solubility
- Write balanced net ionic equations to represent precipitation and neutralization reactions

S'Mores!

- What do we need to make a S'more?

2 Cracker Squares + 1 Marshmallow + 3 Chocolate Squares ⟶ 1 S'more

2 Cq + 1 Mm + 3 Ch ⟶ 1 Cq_2MmCh_3

- The ratio of reactants is 2:1:3 to produce 1 entity of product → Mole Ratio!

Recall!

- Stoichiometry involves calculating the amount of reactants and products in a chemical reaction
- If you know the atoms or ions in the reaction you can use stoichiometry to solve for the rest

$$3KOH + H_3PO_4 \rightarrow K_3PO_4 + 3H_2O$$

- Great news…. Solving stoichiometry problems for solutions uses the same strategies!
  - With the occasional additional steps

Stoichiometry of Solutions

Solve:

1. Write double displacement reactions by their net ionic equations and list given values and molar masses

2. Problems may require you to calculate amount of reactant, given the volume and concentration of the solution

Example 1

19.8g of potassium chloride was dissolved in 100mL of solution. Calculate the concentration (in mol/L) of chloride ions.

Pre-Work: Write balanced net ionic equation and list given values and molar masses

Total ionic: $KCl(s) + H_2O(l) \rightarrow K^+(aq) + Cl^-(aq) + H_2O(l)$

Net ionic: $KCl(s) \rightarrow K^+(aq) + Cl^-(aq)$

m: $m_{KCl} = 19.80g$

M: $M_{KCl} = 74.55g/mol$

Required: Concentration of chloride ions (mol/L)

Example 1

Step 1: Convert **mass to moles**.

Net ionic: $KCl(s) \rightarrow K^+(aq) + Cl^-(aq)$

m: $m_{KCl} = 19.80g$
M: $M_{KCl} = 74.55g/mol$

$n = m / M$

$n_{KCl} = 19.80\ g / 74.55.01\ g/mol = 0.266mol_{KCl}$

Example 1

$$KCl(s) \rightarrow K^+(aq) + Cl^-(aq)$$

m: $m_{KCl} = 19.80g$
M: $M_{KCl} = 74.55g/mol$

Step 2: Convert **moles to moles.** Using the Mole ratio.

Know:

$$n_{KCl} = 0.266mol_{KCl}$$

$$= 0.266mol_{KCl} \times \frac{1\ mol_{Cl}}{1\ mol_{KCl}}$$

$$= 0.266mol_{Cl}$$

Example 1

Step 3: Convert **moles to concentration (mol/L).**

Know:

$n_{Cl} = 0.266mol$

$V_{Solution} = 100mL / 1000L = 0.100L$

$C = n / V$

$= 0.266mol/0.100L$

$= 2.66\ M$

Final Statement: The concentration of of chloride ions when 19.80g of potassium chloride dissolves in 100mL of solution is 2.66M.

Example 2

26.5g of calcium chloride was dissolved in 150mL of solution. Calculate the concentration (in mol/L) of chloride ions.

Pre-Work:

Total ionic: $CaCl_2(s) + H_2O(l) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq) + H_2O(l)$

Net ionic: $CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$

m: $m_{CaCl_2} = 26.5g$

M: $M_{CaCl_2} = 110.98g/mol$

Required: Concentration of chloride ions (mol/L)

Solution

Net ionic: $CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$

m: $m_{CaCl_2} = 26.5g$

M: $M_{CaCl_2} = 110.98g/mol$

Step 1: Convert **mass to moles**.

$26.5g / 110.98g/mol = 0.239\ mol_{CaCl_2}$

Step 2: Convert **moles to moles.**

$$0.239\ mol_{CaCl_2} \times \frac{2\ mol_{Cl}}{1 mol_{CaCl_2}} = 0.478 mol_{Cl}$$

Step 3: Convert **moles to concentration (mol/L).**

$0.478mol / 0.150L = 3.19\ M$

Example 3

Leaves of rhubarb plant contains a relatively high concentration of oxalate ions, $C_2O_4^{2-}$. To determine the percent of oxalate ions, a student measured the mass of some leaves. Then the student ground up the leaves and added excess calcium chloride solution to precipitate calcium oxalate. The student tested 238.6g of leaves. The dried mass of calcium oxalate was 0.556g. What was the mass percent of oxalate in the leaves?

Pre-Work:

Total ionic: $C_2O_4^{2-}(s) + Ca^{2+}(aq) + Cl^-(aq) \rightarrow CaC_2O_4(s) + Cl^-(aq)$

Net ionic: $Ca^{2+}(aq) + C_2O_4^{2-}(aq) \rightarrow CaC_2O_4(s)$

m: $m_{CaC2O4} = 0.556\ g$

M: $M_{CaC2O4} = 128.1 g/mol$

Required: Mass percent of oxalate

Solution

Net ionic: $Ca^{2+}(aq) + C_2O_4^{2-}(aq) \rightarrow CaC_2O_4(s)$

m: $m_{C2O4} = 0.556\,g$

M: $M_{C2O4} = 128.1\,g/mol$ $\quad M_{CaC2O4} = 88.02\,g/mol$

Step 1: Convert **mass to moles**.

$0.556g / 128.1g/mol = 0.00434mol_{C2O4}$

Step 2: Convert **moles to moles.**

$$0.00434\ mol_{C2O4} \times \frac{1\ mol_{Cl}}{1mol_{CaCl2}} = 0.00434mol_{CaC2O4}$$

Step 3: Convert **moles to mass.**

$0.00434mol_{CaC2O4} \times 88.02g/mol = 0.382g_{CaC2O4}$

Solution

$$P_{m/m} = \frac{m_{solute}}{m_{solvent}} \times 100\%$$

$=0.382g/\ 238.6g \times 100\%$

$=0.160\%$

Final Statement: The mass percent of oxalate ions in the leaves is 0.160%

Example 4

Aqueous solutions that silver ions are usually treated with chloride ions to recover silver chloride. What is the minimum volume of 0.25 M magnesium chloride, needed to precipitate all the silver ions in 60 mL of 0.30M silver nitrate?

Pre-Work:

Equation: $MgCl_2\,(aq) + 2AgNO_3\,(aq) \rightarrow 2AgCl\,(s) + Mg(NO_3)_2\,(aq)$

V:	**?**	60mL
C:	0.25M	0.30M

Required: Volume of magnesium chloride

Solution

Equation: $MgCl_2\,(aq) + 2AgNO_3\,(aq) \rightarrow 2AgCl\,(s) + Mg(NO_3)_2\,(aq)$

V:	**?**	60mL
C:	0.25M	0.30M

Step 1: Convert **Concentration to moles**.

$0.06L \times 0.30\ mol/L = 0.018mol_{AgNO3}$

Step 2: Convert **moles to moles.**

$$0.018mol_{AgNO3} \times \frac{1\ mol_{MgCl2}}{\cancel{2mol_{AgNO3}}} = 0.0090mol_{MgCl2}$$

Step 3: Convert **moles to volume**

$0.0090mol_{MgCl2} / 0.25\ mol/L = 0.036L$

Solution

Final Statement: The minimum volume of 0.25 mol?L magnesium chloride that is needed is 0.036L or 36mL

Questions?

Homework

- Discussion questions from your lab (Part A #1-3 , Part B #1 & 4) are due Friday!

- Work on Stoichiometry practice problems

Limiting Reagent

- Determining which of the two reactants is the limiting reactant
  - Using that reactant for further calculations

- In aqueous solutions this usually mean finding the amount of a reactant, given the volume and concentration of the solution