Worked example 2 – Chapter 1

This is an example of how to tackle mole problems where concentrations of reactants are given.

A 35.0 cm^3 sample of a 0.100 mol dm^{-3} potassium chloride solution was reacted with 28.6 cm^3 of a 0.160 mol dm^{-3} silver nitrate solution.

$$AgNO_3(aq) + KCl(aq) \rightarrow AgCl(a) + KNO_3(aq)$$

Calculate the mass of silver chloride precipitated.

Step 1: Work out the number of moles of potassium and silver nitrate. You have been given enough information to work out the number of moles of both reactants, so you should consider the idea that one of them will be a limiting reactant.

number of moles = concentration × volume in dm^3

$$\text{number of moles of KCl} = \frac{35.0}{1000} \times 0.100 = 3.50 \times 10^{-3} \text{ mol}$$

$$\text{number of moles of } AgNO_3 = \frac{28.6}{1000} \times 0.160 = 4.58 \times 10^{-3} \text{ mol}$$

There are fewer moles of KCl, so that is the limiting reactant.

Step 2: The chemical equation tells us that, working with the limiting reactant, 3.50×10^{-3} mol KCl will produce 3.50×10^{-3} mol AgCl.

Step 3: Convert moles of AgCl to mass.

molar mass of AgCl = 143.32 g mol^{-1}

mass of AgCl = number of moles × molar mass

$= 3.50 \times 10^{-3} \times 143.32$

$= 0.502$ g