

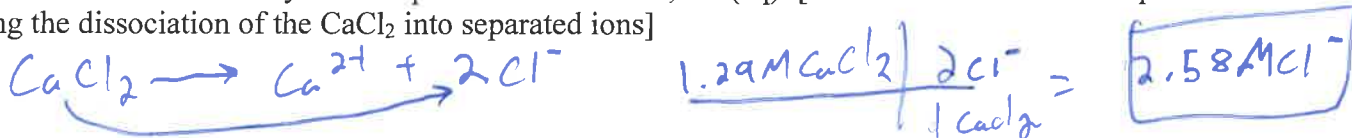
Show all work!!!

1. 12.2-grams of solid calcium chloride,  $\text{CaCl}_2$ , is dissolved in enough water to make 85.0 mL of solution.

(a) Calculate the molarity of the aqueous calcium chloride,  $\text{CaCl}_2$  (aq).

$$\frac{12.2 \text{ g}}{110.98 \text{ g/mol}} \times \frac{1 \text{ mol}}{0.085 \text{ L}} = 1.29 \text{ M CaCl}_2$$

(b) Calculate the molarity of the aqueous chloride ions,  $\text{Cl}^-$  (aq). [Hint: Write a balanced equation showing the dissociation of the  $\text{CaCl}_2$  into separated ions]



2. 25.0 mL of 0.20 M KCl solution is mixed with 15.0 mL of 0.30 M  $\text{Na}_2\text{SO}_4$  solution.

(a) Calculate the molarities of the aqueous KCl (aq) and  $\text{Na}_2\text{SO}_4$  (aq).

$$\frac{0.2 \text{ mol KCl}}{1 \text{ L}} \times 0.025 \text{ L} = 0.005 \text{ mol KCl}$$

$$\frac{0.3 \text{ mol Na}_2\text{SO}_4}{1 \text{ L}} \times 0.015 \text{ L} = 0.0045 \text{ mol Na}_2\text{SO}_4$$

$$\frac{0.005 \text{ mol KCl}}{0.040 \text{ L}} = 0.125 \text{ M KCl}$$

$$\frac{0.0045 \text{ mol Na}_2\text{SO}_4}{0.040 \text{ L}} = 0.1125 \text{ M Na}_2\text{SO}_4$$

$2\text{KCl} + \text{Na}_2\text{SO}_4 \rightarrow \text{K}_2\text{SO}_4 + 2\text{NaCl}$

(b) Calculate the molarity of the aqueous sodium ions,  $\text{Na}^+$  (aq).

$$1 \text{ mol Na}_2\text{SO}_4 = 2 \text{ mol Na}^+ \rightarrow 2 \times 0.1125 \text{ M} = 0.225 \text{ M Na}^+$$

3. 15.0 mL of 0.50 M NaOH solution is mixed with 25.0 mL of 0.35 M NaCl solution. Calculate the molarity of the sodium ion,  $\text{Na}^+$ , in solution. [Hint: begin this problem by finding the moles of  $\text{Na}^+$  ion in each of the original solutions]

$$\text{NaOH} + \text{NaCl} \rightarrow 2\text{Na}^+ + \text{OH}^- + \text{Cl}^-$$

$$\frac{0.5 \text{ mol}}{1 \text{ L}} \times 0.015 \text{ L} = 0.0075 \text{ mol NaOH}$$

$$\frac{0.35 \text{ mol}}{1 \text{ L}} \times 0.025 \text{ L} = 0.00875 \text{ mol NaCl}$$

$$0.0075 \text{ mol NaOH} + 0.00875 \text{ mol NaCl} = 0.01625 \text{ mol Na}^+$$

$$\frac{0.01625 \text{ mol Na}^+}{(0.015 + 0.025) \text{ L}} = 0.4125 \text{ M Na}^+$$

4. 30.0 mL of 1.0 M HCl is reacted with excess zinc metal.



(a) Calculate the moles of HCl in the 30.0 mL of solution.

$$\frac{1 \text{ mol}}{1 \text{ L}} \times 0.03 \text{ L} = 0.03 \text{ mol HCl}$$

(b) Calculate the moles of hydrogen gas and zinc chloride produced in the reaction.

$$\frac{0.03 \text{ mol HCl}}{2 \text{ mol HCl}} \times \frac{1 \text{ mol H}_2}{1} = 0.015 \text{ mol H}_2$$

$$\frac{0.03 \text{ mol HCl}}{2 \text{ mol HCl}} \times \frac{1 \text{ mol ZnCl}_2}{1} = 0.015 \text{ mol ZnCl}_2$$

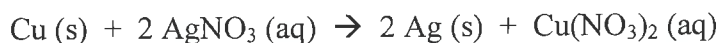
(c) How many liters of hydrogen gas does your answer in (a) represent if conditions are at STP?

$$\frac{0.015 \text{ mol}}{1 \text{ mol}} \times 22.4 \text{ L} = 0.336 \text{ L}$$

(d) Calculate the molarity of the aqueous zinc chloride,  $\text{ZnCl}_2$  (aq), produced.

$$\frac{0.015}{0.336} = 0.045 \text{ M ZnCl}_2$$

5. A piece of copper wire is immersed in 100.0 mL of 0.100 M silver nitrate,  $\text{AgNO}_3$ . Answer the following questions if the copper wire is the excess reactant.



(a) Calculate the moles of  $\text{AgNO}_3$  in the beaker before the copper is added.

$$\frac{0.100 \text{ mol}}{\text{L}} \times 0.1 \text{ L} = 0.01 \text{ mol AgNO}_3$$

(b) Calculate the moles of silver and copper (II) nitrate produced in this reaction. [Hint: If the copper wire is assumed to be excess, then the  $\text{AgNO}_3$  must be limiting!]

$$\frac{0.01 \text{ mol AgNO}_3}{2 \text{ mol AgNO}_3} \times 2 \text{ mol Ag} = 0.01 \text{ mol Ag}$$

$$\frac{0.01 \text{ mol AgNO}_3}{2 \text{ mol AgNO}_3} \times 1 \text{ mol Cu(NO}_3)_2 = 0.005 \text{ mol Cu(NO}_3)_2$$

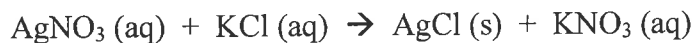
(c) Calculate the grams of silver produced.

$$\frac{0.01 \text{ mol Ag}}{1 \text{ mol}} \times 107.87 \text{ g} = 1.08 \text{ g Ag}$$

(d) Calculate the molarity of the copper (II) nitrate solution produced in this reaction.

$$\frac{0.005 \text{ mol Cu(NO}_3)_2}{0.1 \text{ L}} = 0.05 \text{ M Cu(NO}_3)_2$$

6. 15.0 mL of 0.75 M  $\text{AgNO}_3$  is added to a beaker containing 120.0 mL of 0.10 M  $\text{KCl}$ . Upon mixing the  $\text{AgNO}_3$  and  $\text{KCl}$  undergo a double replacement reaction, precipitating silver chloride,  $\text{AgCl (s)}$ , and producing aqueous potassium nitrate,  $\text{KNO}_3 \text{ (aq)}$ .



(a) Calculate the moles of  $\text{AgNO}_3$  and  $\text{KCl}$  that were initially mixed together.

$$\frac{0.75 \text{ mol AgNO}_3}{\text{L}} \times 0.015 \text{ L} = 0.01125 \text{ mol AgNO}_3$$

$$\frac{0.1 \text{ mol KCl}}{\text{L}} \times 0.120 \text{ L} = 0.012 \text{ mol KCl}$$

(b) Determine the limiting reagent

$$0.01125 \text{ mol AgNO}_3 = \text{limiting since it is less!}$$

(c) Calculate the grams of  $\text{AgCl}$  that precipitated.

$$\frac{0.01125 \text{ mol AgCl}}{1 \text{ mol}} \times 143.32 \text{ g AgCl} = 1.61 \text{ g AgCl}$$