

Chapter 4: Reactions in Aqueous Solution

4. Refer to Section 4.1 and Example 4.1.

- a. Convert from mL to L, then use molarity to convert from L to moles.

$$45.6 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.450 \text{ mol. K}_2\text{CO}_3}{1 \text{ L}} = 0.0205 \text{ mol. K}_2\text{CO}_3$$

- b. Use molarity to convert from moles to L, then convert L to mL.

$$0.800 \text{ mol. K}_2\text{CO}_3 \times \frac{1 \text{ L}}{0.450 \text{ mol. K}_2\text{CO}_3} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 1.78 \times 10^3 \text{ mL}$$

- c. Calculate moles K_2CO_3 present in the 0.450 M solution and the moles needed to make 1 M solution. The difference is the moles needed to bring the solution up to 1 M.

$$2.00 \text{ L} \times \frac{0.450 \text{ mol. K}_2\text{CO}_3}{1 \text{ L}} = 0.900 \text{ mol. K}_2\text{CO}_3$$

$$2.00 \text{ L} \times \frac{1.000 \text{ mol. K}_2\text{CO}_3}{1 \text{ L}} = 2.00 \text{ mol. K}_2\text{CO}_3$$

$$\text{moles needed} = 2.00 \text{ mol.} - 0.900 \text{ mol.} = 1.10 \text{ mol.}$$

$$1.10 \text{ mol. K}_2\text{CO}_3 \times \frac{138.21 \text{ g K}_2\text{CO}_3}{1 \text{ mol. K}_2\text{CO}_3} = 152 \text{ g K}_2\text{CO}_3$$

- d. Calculate moles in 50.0 mL of solution, then recalculate molarity with the new volume.

$$50.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.450 \text{ mol. K}_2\text{CO}_3}{1 \text{ L}} = 0.0225 \text{ mol. K}_2\text{CO}_3$$

$$M = \frac{0.0225 \text{ mol. K}_2\text{CO}_3}{0.125 \text{ L}} = 0.180 \text{ M K}_2\text{CO}_3$$

Alternatively, you could use the equation: $(M_{\text{init}})(V_{\text{init}}) = (M_{\text{final}})(V_{\text{final}})$.

$$(0.450 \text{ M})(50.0 \text{ mL}) = (M_{\text{final}})(125 \text{ mL})$$

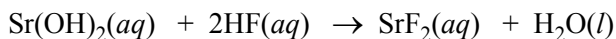
$$(M_{\text{final}}) = 0.180 \text{ M}$$

12. Refer to Sections 2.6 and 4.2, Figures 4.2 and 4.3, and Example 4.4.

Recall that soluble salts ionize when dissolved. Write the reactions for the ionizations. Look at the resulting ions. If there are pairs that would result in insoluble salts, these salts would form and precipitate from solution.

- a. Ions present: Ca^{2+} , SO_4^{2-} , Na^+ , CO_3^{2-}
Possible precipitates: CaCO_3 , Na_2SO_4
 CaCO_3 would precipitate from solution (Na_2SO_4 is soluble and remains in solution).
 $\text{Ca}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightarrow \text{CaCO}_3(\text{s})$ (net ionic equation)
- b. Ions present: Fe^{3+} , SO_4^{2-} , Ba^{2+} , OH^-
Possible precipitates: BaSO_4 , $\text{Fe}(\text{OH})_3$
 BaSO_4 and $\text{Fe}(\text{OH})_3$ would both precipitate from solution.
 $\text{Fe}^{3+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Ba}^{2+}(\text{aq}) + 3\text{OH}^-(\text{aq}) \rightarrow \text{BaSO}_4(\text{s}) + \text{Fe}(\text{OH})_3(\text{s})$

32. Refer to Section 4.3 and Example 4.7.



$$25.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.275 \text{ mol. HF}}{1 \text{ L}} \times \frac{1 \text{ mol. Sr}(\text{OH})_2}{2 \text{ mol. HF}} \\ \times \frac{1 \text{ L}}{0.285 \text{ mol. Sr}(\text{OH})_2} = 0.0121 \text{ L Sr}(\text{OH})_2 \text{ solution}$$

60. Refer to Sections 2.6 and 4.4 and Example 4.10.

Rewrite the reaction, deleting those species that do not change oxidation number. Balance the two half reactions. Balance the element being oxidized or reduced and then balance the oxidation number by adding electrons. Balance the charge by adding H^+ , and H_2O to balance H and O. If the number of electrons in the balanced half reactions are not equal, multiply the equations through with the appropriate coefficient. Combine the two half reactions and cancel those species which appear on both sides of the reaction.

