

1. You have a 0.50 mM solution of sodium phosphate.

A) Determine the concentration of sodium ions and phosphate ions in this solution.

Sodium phosphate is Na_3PO_4 , so $1 \text{ mol Na}_3\text{PO}_4 \rightarrow 3 \text{ mol Na}^+ + 1 \text{ mol PO}_4^{3-}$. Therefore:

$$[\text{Na}^+] = 0.50 \text{ mM Na}_3\text{PO}_4 \times \frac{1 \text{ M}}{1000 \text{ mM}} \times \frac{3 \text{ mol Na}^+}{1 \text{ mol Na}_3\text{PO}_4} = 1.5 \times 10^{-3} \text{ M}$$

$$[\text{PO}_4^{3-}] = 0.50 \text{ mM Na}_3\text{PO}_4 \times \frac{1 \text{ M}}{1000 \text{ mM}} \times \frac{1 \text{ mol PO}_4^{3-}}{1 \text{ mol Na}_3\text{PO}_4} = 5.0 \times 10^{-4} \text{ M}$$

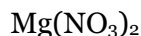
B) If you want to dilute 1.0 mL of the 0.50 mM sodium phosphate solution to 5.0 μM , how much water must be added to the original solution?

We want to convert the 0.50 mM ($5.0 \times 10^{-4} \text{ M}$) solution to 5.0 μM ($5.0 \times 10^{-6} \text{ M}$). We can set up the following expression, which tells us what volume of the 5.0 μM solution would contain the same number of moles as 1.0 mL of the 0.50 mM solution.

$$\begin{aligned} M_1 V_1 &= M_2 V_2 \\ (5.0 \times 10^{-4} \text{ M})(0.0010 \text{ L}) &= (5.0 \times 10^{-6} \text{ M})V_2 \\ V_2 &= 0.10 \text{ L (or 100 mL)} \end{aligned}$$

To get a V_2 of 100 mL, we need to add: $100 \text{ mL} - 1.0 \text{ mL} = 99 \text{ mL water}$.

2. Consider the following four aqueous solutions:

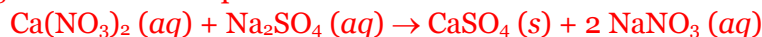


A) Which of the aqueous solutions would produce a precipitate when added to an aqueous solution of sodium sulfate?

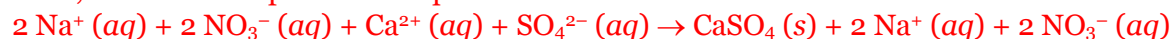


B) Write the balanced net ionic equation for the precipitation reaction in part A.

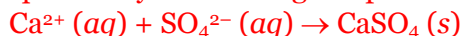
Start by writing the molecular equation:



Then, write the complete ionic equation:

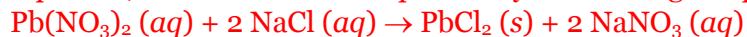


Finally, write the net ionic equation by eliminating all spectator ions:



3. Determine the mass of solid precipitate formed when an excess lead(II) nitrate solution is mixed with 0.0800 mol sodium chloride.

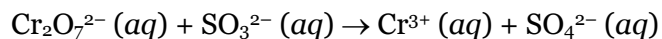
Write the molecular equation, then the net ionic equation by eliminating all spectator ions:



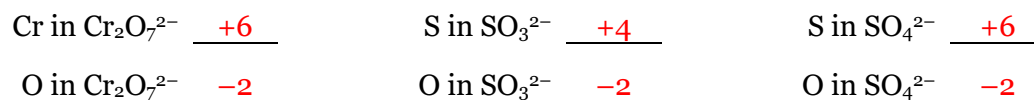
Because the Pb^{2+} is in excess, the Cl^- is limiting. Hence,

$$0.0800 \text{ mol NaCl} \times \frac{1 \text{ mol Cl}^-}{1 \text{ mol NaCl}} \times \frac{1 \text{ mol PbCl}_2}{2 \text{ mol Cl}^-} \times \frac{278.1 \text{ g}}{1 \text{ mol PbCl}_2} = 11.1 \text{ g PbCl}_2$$

4. Consider the following unbalanced reaction:



A) Determine the oxidation states for the specified atoms.



B) Balance the above reaction using the half-reaction method in acidic aqueous solution.

1. Set up the two half-reactions: $\text{SO}_3^{2-} \rightarrow \text{SO}_4^{2-}$ (reduction: $\text{S}^{4+} \rightarrow \text{S}^{6+}$)
 $\text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}^{3+}$ (oxidation: $\text{Cr}^{6+} \rightarrow \text{Cr}^{3+}$)

2. Balance elements (except O & H): $\text{SO}_3^{2-} \rightarrow \text{SO}_4^{2-}$
 $\text{Cr}_2\text{O}_7^{2-} \rightarrow 2 \text{Cr}^{3+}$

3. Balance O atoms with H_2O : $\text{SO}_3^{2-} + \text{H}_2\text{O} \rightarrow \text{SO}_4^{2-}$
 $\text{Cr}_2\text{O}_7^{2-} \rightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O}$

4. Balance H atoms with H^+ : $\text{SO}_3^{2-} + \text{H}_2\text{O} \rightarrow \text{SO}_4^{2-} + 2 \text{H}^+$
 $\text{Cr}_2\text{O}_7^{2-} + 14 \text{H}^+ \rightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O}$

5. Balance charges with e^- : $\text{SO}_3^{2-} + \text{H}_2\text{O} \rightarrow \text{SO}_4^{2-} + 2 \text{H}^+ + 2 e^-$
 $\text{Cr}_2\text{O}_7^{2-} + 14 \text{H}^+ + 6 e^- \rightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O}$

5. Balance electrons: $3 \times [\text{SO}_3^{2-} + \text{H}_2\text{O} \rightarrow \text{SO}_4^{2-} + 2 \text{H}^+ + 2 e^-]$
 $\text{Cr}_2\text{O}_7^{2-} + 14 \text{H}^+ + 6 e^- \rightarrow 2 \text{Cr}^{3+} + 7 \text{H}_2\text{O}$

6. Sum up the two half-reactions: $\text{Cr}_2\text{O}_7^{2-} + 3 \text{SO}_3^{2-} + 8 \text{H}^+ \rightarrow 2 \text{Cr}^{3+} + 3 \text{SO}_4^{2-} + 4 \text{H}_2\text{O}$

5. A titration is performed where 48.0 mL of 1.00 M HCl is needed to react completely with 20.0 mL of a LiOH solution with unknown concentration.

What is the concentration of the unknown LiOH solution?

First, determine how many moles of HCl or H^+ ions (mole ratio between HCl: H^+ is 1:1) were required to reach the end-point:

$$n_{\text{HCl}} = 48.0 \text{ mL HCl} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1.00 \text{ mol HCl}}{1 \text{ L}} = 0.0480 \text{ mol HCl} \quad n_{\text{H}^+} = 0.0480 \text{ mol H}^+$$

Second, recognize that the end-point occurs when $n_{\text{acid}} = n_{\text{base}}$, which means we have 0.0480 mol OH^- at the end-point. All of these OH^- ions come from solution B (diluted LiOH), which means we must have started with 0.0480 mol LiOH since the LiOH: OH^- mole ratio is 1:1.

Finally, we can now find the concentration of the LiOH solution:

$$M_{\text{LiOH}} = \frac{0.0480 \text{ mol LiOH}}{20.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}}} = 2.40 \text{ M}$$

Note: An alternative solution would be to start with this equivalent expression:

$$\left(1.00 \frac{\text{mol HCl}}{1 \text{ L}}\right) \left(48.0 \text{ mL HCl} \times \frac{1 \text{ L}}{1000 \text{ mL}}\right) = M_{\text{base}} \left(20.0 \text{ mL HCl} \times \frac{1 \text{ L}}{1000 \text{ mL}}\right)$$

$$M_{\text{base}} = 2.40 \text{ M}$$