Aqueous Reactions and Solution Stoichiometry

Visualizing Concepts

4.1 Analyze. Correlate the formula of the solute with the charged spheres in the diagrams.

Plan. Determine the electrolyte properties of the solute and the relative number of cations, anions, or neutral molecules produced when the solute dissolves.

Solve. Li$_2$SO$_4$ is a strong electrolyte, a soluble ionic solid that dissociates into separate Li$^+$ and SO$_4^{2-}$ when it dissolves in water. There are twice as many Li$^+$ cations as SO$_4^{2-}$ anions. Diagram (c) represents the aqueous solution of a 2:1 electrolyte.

4.3 Analyze/Plan. Correlate the neutral molecules, cations, and anions in the diagrams with the definitions of strong, weak, and nonelectrolytes. Solve.

(a) AX is a nonelectrolyte, because no ions form when the molecules dissolve.

(b) AY is a weak electrolyte because a few molecules ionize when they dissolve, but most do not.

(c) AZ is a strong electrolyte because all molecules break up into ions when they dissolve.

4.5 Analyze. From the names and/or formulas of 3 substances determine their electrolyte and solubilities properties.

Plan. Determine whether the substance is molecular or ionic. If it is molecular, is it a weak acid or base and thus a weak electrolyte, or a nonelectrolyte? If it is ionic, is it soluble?

Solve. Glucose is a molecular compound that is neither a weak acid nor a weak base. It is a nonelectrolyte that dissolves to produce a nonconducting solution; it is solid C. NaOH is ionic; the anion is OH$^-$. According to Table 4.1, most hydroxides are insoluble, but NaOH is one of the soluble ones. NaOH is a strong electrolyte that dissolves to form a conducting solution; it is solid A. AgBr is ionic; the anion is Br$^-$. According to Table 4.1, most bromides are soluble, but AgBr is one of the insoluble ones; it is solid B.

Check. We know by elimination that AgBr is solid B, which we verified by solubility rules.

4.7 Analyze. Given the formulas of some ions, determine whether these ions ever form precipitates in aqueous solution. Plan. Use Table 4.1 to determine if the given ions can form precipitates. If not, they will always be spectator ions. Solve.
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(a) Cl\(^-\) can form precipitates with Ag\(^+\), Hg\(_2\)\(^{2+}\), Pb\(^{2+}\).
(b) NO\(_3\)\(^-\) never forms precipitates, so it is always a spectator.
(c) NH\(_4\)\(^+\) never forms precipitates, so it is always a spectator.
(d) S\(^2-\) usually forms precipitates.
(e) SO\(_4\)\(^{2-}\) usually forms precipitates.

Check. NH\(_4\)\(^+\) is a soluble exception for sulfides, phosphates, and carbonates, which usually form precipitates, so all rules indicate that it is a perpetual spectator.

4.9 In a redox reaction, one reactant loses electrons and a different reactant gains electrons; electrons are transferred. Acids ionize in aqueous solution to produce (donate) hydrogen ions (H\(^+\), protons). Bases are substances that react with or accept protons (H\(^+\)). In an acid-base reaction, protons are transferred from an acid to a base. We characterize redox reactions by tracking electron transfer using oxidation numbers. We characterize acid-base reactions by tracking H\(^+\) (proton) transfer via molecular formulas of reactants and products.

Electrolytes

4.11 No. Electrolyte solutions conduct electricity because the dissolved ions carry charge through the solution (from one electrode to the other).

4.13 Although H\(_2\)O molecules are electrically neutral, there is an unequal distribution of electrons throughout the molecule. There are more electrons near O and fewer near H, giving the O end of the molecule a partial negative charge and the H end of the molecule a partial positive charge. Ionic compounds are composed of positively and negatively charged ions. The partially positive ends of H\(_2\)O molecules are attracted to the negative ions (anions) in the solid, while the partially negative ends are attracted to the positive ions (cations). Thus, both cations and anions in an ionic solid are surrounded and separated (dissolved) by H\(_2\)O molecules.

4.15 Analyze/Plan. Given the solute formula, determine the separate ions formed upon dissociation. Solve.
(a) ZnCl\(_2\) (aq) \(\rightarrow\) Zn\(^{2+}\) (aq) + 2Cl\(^-\) (aq)
(b) HNO\(_3\) (aq) \(\rightarrow\) H\(^+\) (aq) + NO\(_3\)\(^-\) (aq)
(c) (NH\(_4\)\(_2\))\(_2\)SO\(_4\) (aq) \(\rightarrow\) 2NH\(_4\)\(^+\) (aq) + SO\(_4\)\(^{2-}\) (aq)
(d) Ca(OH\(_2\)) (aq) \(\rightarrow\) Ca\(^{2+}\) (aq) + 2OH\(^-\) (aq)

4.17 Analyze/Plan. Apply the definition of a weak electrolyte to HCOOH.
Solve. When HCOOH dissolves in water, neutral HCOOH molecules, H\(^+\) ions and HCOO\(^-\) ions are all present in the solution. HCOOH (aq) \(\rightleftharpoons\) H\(^+\) (aq) + HCOO\(^-\) (aq)

Precipitation Reactions and Net Ionic Equations

Plan. Follow the guidelines in Table 4.1, in light of the anion present in the compound and notable exceptions to the “rules.” Solve.
4 Aqueous Reactions

(a) $\text{NiCl}_2$: soluble
(b) $\text{Ag}_2\text{S}$: insoluble
(c) $\text{Cs}_2\text{PO}_4$: soluble ($\text{Cs}^+$ is an alkali metal cation)
(d) $\text{SrCO}_3$: insoluble
(e) $\text{PbSO}_4$: insoluble, $\text{Pb}^{2+}$ is an exception to soluble sulfates


Plan. Follow the logic in Sample Exercise 4.3.

Solve. In each reaction, the precipitate is in bold type.

(a) $\text{Na}_2\text{CO}_3(\text{aq}) + 2\text{AgNO}_3(\text{aq}) \rightarrow \text{Ag}_2\text{CO}_3(\text{s}) + 2\text{NaNO}_3(\text{aq})$
(b) No precipitate (all nitrates and most sulfates are soluble).
(c) $\text{FeSO}_4(\text{aq}) + \text{Pb(NO}_3)_2(\text{aq}) \rightarrow \text{PbSO}_4(\text{s}) + \text{Fe(NO}_3)_2(\text{aq})$

4.23 Analyze/Plan. Follow the logic in Sample Exercise 4.4. From the complete ionic equation, identify the ions that don’t change during the reaction; these are the spectator ions.

Solve.

(a) $2\text{Na}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) + \text{Mg}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{MgCO}_3(\text{s}) + 2\text{Na}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq})$
Spectators: $\text{Na}^+, \text{SO}_4^{2-}$
(b) $\text{Pb}^{2+}(\text{aq}) + 2\text{NO}_3^-(\text{aq}) + 2\text{Na}^+(\text{aq}) + \text{S}^{2-}(\text{aq}) \rightarrow \text{PbS}(\text{s}) + 2\text{Na}^+(\text{aq}) + 2\text{NO}_3^-(\text{aq})$
Spectators: $\text{Na}^+, \text{NO}_3^-$
(c) $6\text{NH}_4^+(\text{aq}) + 2\text{PO}_4^{3-}(\text{aq}) + 3\text{Ca}^{2+}(\text{aq}) + 6\text{Cl}^-(\text{aq}) \rightarrow \text{Ca}_2(\text{PO}_4)_2(\text{s}) + 6\text{NH}_4^+(\text{aq}) + 6\text{Cl}^-(\text{aq})$
Spectators: $\text{NH}_4^+, \text{Cl}^-$

4.25 Analyze. Given: reactions of unknown with $\text{HBr}$, $\text{H}_2\text{SO}_4$, $\text{NaOH}$. Find: The unknown contains a single salt. Is $\text{K}^+$ or $\text{Pb}^{2+}$ or $\text{Ba}^{2+}$ present?

Plan. Analyze solubility guidelines for $\text{Br}^-$, $\text{SO}_4^{2-}$ and $\text{OH}^-$ and select the cation that produces a precipitate with each of the anions.

Solve. $\text{K}^+$ forms no precipitates with any of the anions. $\text{BaSO}_4$ is insoluble, but $\text{BaCl}_2$ and $\text{Ba(OH)}_2$ are soluble. Since the unknown forms precipitates with all three anions, it must contain $\text{Pb}^{2+}$.

Check. $\text{PbBr}_2$, $\text{PbSO}_4$, and $\text{Pb(OH)}_2$ are all insoluble according to Table 4.1, so our process of elimination is confirmed by the insolubility of the $\text{Pb}^{2+}$ compounds.

4.27 Analyze. Given: three possible salts in an unknown solution react with $\text{Ba(NO}_3)_2$ and then $\text{NaCl}$. Find: Can the results identify the unknown salt? Do the three possible unknowns give distinctly different results with $\text{Ba(NO}_3)_2$ and $\text{NaCl}$?

Plan. Using Table 4.1, determine whether each of the possible unknowns will form a precipitate with $\text{Ba(NO}_3)_2$ and $\text{NaCl}$. Solve.

<table>
<thead>
<tr>
<th>Compound</th>
<th>$\text{Ba(NO}_3)_2$ result</th>
<th>$\text{NaCl}$ result</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{AgNO}_3(\text{aq})$</td>
<td>no ppt</td>
<td>$\text{AgCl}$ ppt</td>
</tr>
</tbody>
</table>
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<p>| | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>CaCl$_2$(aq)</td>
<td>no ppt</td>
<td>no ppt</td>
</tr>
<tr>
<td>Al$_2$(SO$_4$)$_3$</td>
<td>BaSO$_4$ ppt</td>
<td>no ppt</td>
</tr>
</tbody>
</table>

This sequence of tests would definitively identify the contents of the bottle, because the results for each compound are unique.

Acid-Base Reactions


Plan: See Sample Exercise 4.6. Determine whether solutes are strong or weak acids or bases, or nonelectrolytes. For solutions of equal concentration, strong acids will have greatest concentration of solvated protons. Take varying concentration into consideration when evaluating the same class of solutions.

Solve. LiOH is a strong base, HI is a strong acid, CH$_3$OH is a molecular compound and nonelectrolyte. The strong acid HI will have the greatest concentration of solvated protons.

Check. The solution concentrations weren’t needed to answer the question.

4.31  (a) A monoprotic acid has one ionizable (acidic) H and a diprotic acid has two.

(b) A strong acid is completely ionized in aqueous solution, whereas only a fraction of weak acid molecules are ionized.

(c) An acid is an H$^+$ donor, a substance that increases the concentration of H$^+$ in aqueous solution. A base is an H$^+$ acceptor and thus increases the concentration of OH$^-$ in aqueous solution.

4.33  Strong acids are completely ionized, while weak acids are partially or slightly ionized. As strong acids, HCl, HBr and HI are completely ionized in aqueous solution. They exist as H$^+$ (aq) and Cl$^-$ (aq), H$^+$ (aq) and Br$^-$ (aq) and H$^+$ (aq) and I$^-$ (aq), respectively. There are no neutral HX molecules in these solutions. As a weak acid, HF exists as a mixture of H$^+$ (aq), F$^-$ (aq) and HF(aq). There are at least as many neutral HF molecules as anions or cations. For equal solution concentrations, HCl, HBr and HI will produce a bright light like Figure 4.2(c), while HF will produce a dim light like Figure 4.2(b).

4.35  Analyze. Given: chemical formulas. Find: classify as acid, base, salt; strong, weak, or nonelectrolyte.

Plan. See Table 4.3. Ionic or molecular? Ionic, soluble: OH$^-$, strong base and strong electrolyte; otherwise, salt, strong electrolyte. Molecular: NH$_3$, weak base and weak electrolyte; H-first, acid; strong acid (Table 4.2), strong electrolyte; otherwise weak acid and weak electrolyte.  Solve.

(a) HF: acid, mixture of ions and molecules (weak electrolyte)

(b) CH$_3$CN: none of the above, entirely molecules (nonelectrolyte)

(c) NaClO$_4$: salt, entirely ions (strong electrolyte)

(d) Ba(OH)$_2$: base, entirely ions (strong electrolyte)

Plan. In order to classify as electrolytes, formulas must be identified as acids, bases, or salts as in Solution 4.35. Solve.

(a) \( \text{H}_2\text{SO}_3 \): H first, so acid; not in Table 4.2, so weak acid; therefore, weak electrolyte

(b) \( \text{C}_2\text{H}_5\text{OH} \): not acid, not ionic (no metal cation), contains OH group, but not as anion so not a base; therefore, nonelectrolyte

(c) \( \text{NH}_3 \): common weak base; therefore, weak electrolyte

(d) \( \text{KClO}_3 \): ionic compound, so strong electrolyte

(e) \( \text{Cu(NO}_3)_2 \): ionic compound, so strong electrolyte


(a) \( 2\text{HBr(aq)} + \text{Ca(OH)}_2(\text{aq}) \rightarrow \text{CaBr}_2(\text{aq}) + 2\text{H}_2\text{O(l)} \)

(b) \( \text{Cu(OH)}_2(\text{s}) + 2\text{HClO}_4(\text{aq}) \rightarrow \text{Cu(ClO}_4)_2(\text{aq}) + 2\text{H}_2\text{O(l)} \)

(c) \( \text{Al(OH)}_3(\text{s}) + 3\text{HNO}_3(\text{aq}) \rightarrow \text{Al(NO}_3)_3(\text{aq}) + 3\text{H}_2\text{O(l)} \)


Plan. Write correct chemical formulas for the reactants, complete and balance the metathesis reaction, and identify either \( \text{H}_2\text{S} \) or \( \text{CO}_2 \) products as gases. Solve.

(a) \( \text{CdS(s)} + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CdSO}_4(\text{aq}) + \text{H}_2\text{S(g)} \)

(b) \( \text{MgCO}_3(\text{s}) + 2\text{HClO}_4(\text{aq}) \rightarrow \text{Mg(ClO}_4)_2(\text{aq}) + \text{H}_2\text{O(l)} + \text{CO}_2(\text{g}) \)

4.43 Analyze. Given the formulas or names of reactants, write balanced molecular and net ionic equations for the reactions.

Plan. Write correct chemical formulas for all reactants. Predict products of the neutralization reactions by exchanging ion partners. Balance the complete molecular equation, identify spectator ions by recognizing strong electrolytes, write the corresponding net ionic equation (omitting spectators). Solve.

(a) \( \text{CaCO}_3(\text{s}) + 2\text{HNO}_3(\text{aq}) \rightarrow \text{Ca(NO}_3)_2(\text{aq}) + \text{H}_2\text{O(l)} + \text{CO}_2(\text{g}) \)

(b) \( \text{FeS(s)} + 2\text{HBr(aq)} \rightarrow \text{FeBr}_2(\text{aq}) + \text{H}_2\text{S(g)} \)
Oxidation-Reduction Reactions

4.45  (a) In terms of electron transfer, *oxidation* is the loss of electrons by a substance, and *reduction* is the gain of electrons (LEO says GER).

(b) Relative to oxidation numbers, when a substance is oxidized, its oxidation number increases. When a substance is reduced, its oxidation number decreases.

4.47  *Analyze.* Given the labeled periodic chart, determine which region is most readily oxidized and which is most readily reduced.

*Plan.* Review the definition of oxidation and apply it to the properties of elements in the indicated regions of the chart.

*Solve.* Oxidation is loss of electrons. Elements readily oxidized form positive ions; these are metals. Elements not readily oxidized tend to gain electrons and form negative ions; these are nonmetals. Elements in regions A, B, and C are metals, and their ease of oxidation is shown in Table 4.5. Metals in region A, Na, Mg, K, and Ca are most easily oxidized. Elements in region D are nonmetals and are least easily oxidized.

4.49  *Analyze.* Given the chemical formula of a substance, determine the oxidation number of a particular element in the substance.

*Plan.* Follow the logic in Sample Exercise 4.8. *Solve.*

(a) +4   (b) +4   (c) +7   (d) +1   (e) 0   (f) −1 (O$_2^{2−}$ is peroxide ion)

4.51  *Analyze.* Given: chemical reaction. Find: element oxidized or reduced. *Plan.* Assign oxidation numbers to all species. The element whose oxidation number becomes more positive is oxidized; the one whose oxidation number decreases is reduced. *Solve.*

(a) $\text{N}_2(g) [\text{N, 0}] \rightarrow 2\text{NH}_3(g) [\text{N, −3}], \text{N is reduced}; 3\text{H}_2(g) [\text{H, 0}] \rightarrow 2\text{NH}_3(g) [\text{H, +1}], \text{H is oxidized.}$

(b) $\text{Fe}^{2+} \rightarrow \text{Fe}, \text{Fe is oxidized}; \text{Al} \rightarrow \text{Al}^{3+}, \text{Al is oxidized}$

(c) $\text{Cl}_2 \rightarrow 2\text{Cl}, \text{Cl is reduced}; 2\text{I}^- \rightarrow \text{I}_2, \text{I is oxidized}$

(d) $\text{S}^{2−} \rightarrow \text{SO}_4^{2−}(\text{S, +6}), \text{S is oxidized}; \text{H}_2\text{O}_2 (\text{O, −1}) \rightarrow \text{H}_2\text{O} (\text{O, −2}); \text{O is reduced}$


*Plan.* Metals oxidized by $\text{H}^+$ form cations. Predict products by exchanging cations and balance. The anions are the spectator ions and do not appear in the net ionic equations. *Solve.*

(a) $\text{Mn}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{MnSO}_4(aq) + \text{H}_2(g); \text{Mn}(s) + 2\text{H}^+(aq) \rightarrow \text{Mn}^{2+}(aq) + \text{H}_2(g)$

Products with the metal in a higher oxidation state are possible, depending on reaction conditions and acid concentration.

(b) $2\text{Cr}(s) + 6\text{HBr}(aq) \rightarrow 2\text{CrBr}_3(aq) + 3\text{H}_2(g); 2\text{Cr}(s) + 6\text{H}^+(aq) \rightarrow 2\text{Cr}^{3+}(aq) + 3\text{H}_2(g)$

(c) $\text{Sn}(s) + 2\text{HCl}(aq) \rightarrow \text{SnCl}_2(aq) + \text{H}_2(g); \text{Sn}(s) + 2\text{H}^+(aq) \rightarrow \text{Sn}^{2+}(aq) + \text{H}_2(g)$

(d) $2\text{Al}(s) + 6\text{HCOOH}(aq) \rightarrow 2\text{Al}($$	ext{HCOO})_3(aq) + 3\text{H}_2(g);$
4 Aqueous Reactions

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2Al(s) + 6HCOOH(aq) → 2Al^{3+}(aq) + 6HCOO^-(aq) + 3H_2(g)


Plan. Use Table 4.5. If the metal is above the aqueous solution, reaction will occur; if the aqueous solution is higher, NR. If reaction occurs, predict products by exchanging cations (a metal ion or H^+), then balance the equation. Solve.

(a) Fe(s) + Cu(NO_3)_2(aq) → Fe(NO_3)_2(aq) + Cu(s)
(b) Zn(s) + MgSO_4(aq) → NR
(c) Sn(s) + 2HBr(aq) → SnBr_2(aq) + H_2(g)
(d) H_2(g) + NiCl_2(aq) → NR
(e) 2Al(s) + 3CoSO_4(aq) → Al_2(SO_4)_3(aq) + 3Co(s)

4.57 (a) i. Zn(s) + Cd^{2+}(aq) → Cd(s) + Zn^{2+}(aq)
ii. Cd(s) + Ni^{2+}(aq) → Ni(s) + Cd^{2+}(aq)
(b) According to Table 4.5, the most active metals are most easily oxidized, and Zn is more active than Ni. Observation (i) indicates that Cd is less active than Zn; observation (ii) indicates that Cd is more active than Ni. Cd is between Zn and Ni on the activity series.

(c) Place an iron strip in CdCl_2(aq). If Cd(s) is deposited, Cd is less active than Fe; if there is no reaction, Cd is more active than Fe. Do the same test with Co if Cd is less active than Fe or with Cr if Cd is more active than Fe.

Solution Composition; Molarity

4.59 (a) Concentration is an intensive property; it is the ratio of the amount of solute present in a certain quantity of solvent or solution. This ratio remains constant regardless of how much solution is present.
(b) The term 0.50 mol HCl defines an amount (~18 g) of the pure substance HCl. The term 0.50 M HCl is a ratio; it indicates that there are 0.50 mol of HCl solute in 1.0 liter of solution. This same ratio of moles solute to solution volume is present regardless of the volume of solution under consideration.

4.61 Analyze/Plan. Follow the logic in Sample Exercises 4.11 and 4.13. Solve.

(a) \[ M = \frac{\text{mol solute}}{\text{L solution}} = \frac{0.025\text{mol NH}_4\text{Cl}}{500\text{mL}} \times \frac{1000\text{mL}}{1\text{L}} = 0.0500\text{M NH}_4\text{Cl} \]
Check. \((0.025 \times 2) \approx 0.050 \text{M}\)

(b) \[ \text{mol} = M \times L = \frac{2.5\text{mol HNO}_3}{1\text{L}} \times 0.0500\text{L} = 0.125\text{mol HNO}_3 \]
Check. \((2.5 \times 0.05) \approx 0.125 \text{mol}\)

(c) \[ L = \frac{\text{mol}}{M} = \frac{0.27\text{mol KOH}}{1.50\text{mol KOH/L}} = 0.188 \text{ or } 183\text{mL of 1.50M KOH} \]
4.63 **Analyze.** Given molarity, \( M \), and volume, \( L \), find mass of \( \text{Na}^+ \text{(aq)} \) in the blood.

**Plan.** Calculate moles \( \text{Na}^+ \text{(aq)} \) using the definition of molarity: \( M = \frac{\text{mol}}{L} \); \( \text{mol} = M \times L \).

Calculate mass \( \text{Na}^+ \text{(aq)} \) using the definition moles: \( \text{mol} = g / \text{MM} \); \( g = \text{mol} \times \text{MM} \). (MM is the symbol for molar mass in this manual.)

\[
\text{Solve.} \quad \frac{0.135\text{mol}}{L} \times 5.0L \times \frac{23.0g}{\text{mol Na}^+} = 1552.5 \times 16g \text{ Na}^+ \text{(aq)}
\]

**Check.** Since there are more than 0.1 mol/L and we have 5.0 L, there should be more than half a mol (11.5 g) of \( \text{Na}^+ \). The calculation agrees with this estimate.

4.65 **Analyze.** Given: g alcohol/100 mL blood; molecular formula of alcohol. Find: molarity (mol/L) of alcohol. **Plan.** Use the molar mass (MM) of alcohol to change (g/100 mL) to (mol/100 mL) then mL to L.

**Solve.** \( \text{MM of alcohol} = 2(12.01) + 6(1.1008) + 1(16.00) = 46.07 \text{ g alcohol/mol} \)

\[
\text{BAC} = \frac{0.01\text{g alcohol}}{100\text{mL blood}} \times \frac{1\text{mol alcohol}}{46.08\text{g alcohol}} \times \frac{1000\text{mL}}{1L} = 0.0174 = 0.02 \text{ M alcohol}
\]

4.67 **Plan.** Proceed as in Sample Exercises 4.11 and 4.13.

\[
M = \frac{\text{mol}}{L} \text{; mol} = \frac{g}{M} \quad \text{(MM is the symbol for molar mass in this manual.)}
\]

**Solve.**

(a) \[
\frac{0.175\text{M KBr}}{1L} \times 0.25L \times \frac{119.0\text{g KBr}}{1\text{mol KBr}} = 5.21\text{g KBr}
\]

**Check.** \( (0.25 \times 120) = 30; 30 \times 0.18 = 5.4 \text{ g KBr} \)

(b) \[
1475\text{g Ca(NO}_3\text{)_2} \times \frac{1\text{mol Ca(NO}_3\text{)_2}}{164.09\text{g Ca(NO}_3\text{)_2}} \times \frac{1}{1.379L} = 0.0653\text{M Ca(NO}_3\text{)_2}
\]

**Check.** \( (15/1.5) = 10; 10/160 = 5/80 \approx 0.06 \text{ M Ca(NO}_3\text{)_2} \)

(c) \[
2.5\text{g Na}_3\text{PO}_4 \times \frac{1\text{mol Na}_3\text{PO}_4}{169.9\text{g Na}_3\text{PO}_4} \times \frac{1L}{1.50\text{mol Na}_3\text{PO}_4} \times \frac{1000\text{mL}}{1L} = 102\text{mL solution}
\]

**Check.** \( [25/(160 \times 1.5)] \approx 2.5/240 \approx 1/100 \approx 0.01 \text{ L = 10 mL} \)

4.69 **Analyze.** Given: formula and concentration of each solute. Find: concentration of \( \text{K}^+ \) in each solution. **Plan.** Note mol \( \text{K}^+ \)/mol solute and compare concentrations or total moles.

**Solve.**

(a) \( \text{KCl} \rightarrow \text{K}^+ + \text{Cl}^- \); \( 0.20 \text{ M KCl} = 0.20 \text{ M K}^+ \)
4 Aqueous Reactions

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\[ K_2\text{Cr}_4O_4 \rightarrow 2 K^+ + \text{CrO}_4^{2-}; \quad 0.15 \, M \, K_2\text{Cr}_4O_4 = 0.30 \, M \, K^+ \]

\[ K_3\text{PO}_4 \rightarrow 3 K^+ + \text{PO}_4^{3-}; \quad 0.080 \, M \, K_3\text{PO}_4 = 0.24 \, M \, K^+ \]

0.15 \, M \, K_2\text{Cr}_4O_4 \text{ has the highest } K^+ \text{ concentration.}

(b) \[ K_2\text{Cr}_4O_4: \quad 0.30 \, M \, K^+ \times 0.0300 \, L = 0.0090 \, \text{mol} \, K^+ \]

\[ K_3\text{PO}_4: \quad 0.24 \, M \, K^+ \times 0.0250 \, L = 0.0060 \, \text{mol} \, K^+ \]

30.0 mL of 0.15 \, M \, K_2\text{Cr}_4O_4 \text{ has more } K^+ \text{ ions.}

4.71 Analyze. Given: molecular formula and solution molarity. Find: concentration (M) of each ion.

Plan. Follow the logic in Sample Exercise 4.12.

Solve.

(a) \[ \text{NaNO}_3 \rightarrow \text{Na}^+ + \text{NO}_3^-, \quad 0.25 \, M \, \text{NaNO}_3 = 0.25 \, M \, \text{Na}^+, \quad 0.25 \, M \, \text{NO}_3^- \]

(b) \[ \text{MgSO}_4 \rightarrow \text{Mg}^{2+}, \, \text{SO}_4^{2-}, \quad 1.3 \times 10^{-2} \, M \, \text{MgSO}_4 = 1.3 \times 10^{-2} \, M \, \text{Mg}^{2+}, \quad 1.3 \times 10^{-2} \, M \, \text{SO}_4^{2-} \]

(c) \[ \text{C}_6\text{H}_12\text{O}_6 \rightarrow \text{C}_6\text{H}_12\text{O}_6 \text{ (molecular solute); } \quad 0.0150 \, M \, \text{C}_6\text{H}_12\text{O}_6 = 0.0150 \, M \, \text{C}_6\text{H}_12\text{O}_6 \]

(d) Plan. There is no reaction between \text{NaCl} and \text{(NH}_4\text{)}_2\text{CO}_3, \text{so this is just a dilution problem, } M_1V_1 = M_2V_2. \text{ Then account for ion stoichiometry.}

Solve. 45.0 mL + 65.0 mL = 110.0 mL total volume

\[ \frac{0.272 M \, \text{NaCl} \times 45.0 \, \text{mL}}{110.0 \, \text{mL}} = 0.111 \, M \, \text{NaCl}; \quad 0.111 \, M \, \text{Na}^+, \quad 0.111 \, M \, \text{Cl}^- \]

\[ \frac{0.247 M \, (\text{NH}_4)_2\text{CO}_3 \times 65.0 \, \text{mL}}{110.0 \, \text{mL}} = 0.0146 \, M \, (\text{NH}_4)_2\text{CO}_3 \]

\[ 2 \times (0.0146 \, M) = 0.0292 \, M \, \text{NH}_4^+; \quad 0.0146 \, M \, \text{CO}_3^{2-} \]

Check. By adding the two solutions (with no common ions or chemical reaction), we have approximately doubled the solution volume, and reduced the concentration of each ion by approximately a factor of two.

4.73 Analyze/Plan. Follow the logic of Sample Exercise 4.14.

Solve.

(a) \[ V_1 = \frac{M_2V_2}{M_1}; \quad \frac{0.250 M \, \text{NH}_3 \times 1000.0 \, \text{mL}}{14.8 M \, \text{NH}_3} = 16.89 \, \text{mL} \quad 14.8 M \, \text{NH}_3 \]

Check. 250/15 = 15 \, M

(b) \[ M_2 = M_1V_1 / V_2; \quad \frac{14.8 M \, \text{NH}_3 \times 100.0 \, \text{mL}}{500 \, \text{mL}} = 0.296 M \, \text{NH}_3 \]

Check. 150/500 = 0.30 \, M
4.75 (a) Plan/Solve. Follow the logic in Sample Exercise 4.13. The number of moles of sucrose needed is 0.250 mol/L x 0.250 L = 0.0625 mol.

Weigh out 0.062 mol C₁₂H₂₂O₁₁ x \frac{342.2 g C₁₂H₂₂O₁₁}{1 mol C₁₂H₂₂O₁₁} = 214 g C₁₂H₂₂O₁₁

Add this amount of solid to a 250 mL volumetric flask, dissolve in a small volume of water, and add water to the mark on the neck of the flask. Agitate thoroughly to ensure total mixing.

(b) Plan/Solve. Follow the logic in Sample Exercise 4.14. Calculate the moles of solute present in the final 400.0 mL of 0.100 M C₁₂H₂₂O₁₁ solution:

\text{moles C₁₂H₂₂O₁₁} = \frac{0.100 \text{ mol C₁₂H₂₂O₁₁}}{1 \text{ L}} x 0.350 \text{ L} = 0.035 \text{ mol C₁₂H₂₂O₁₁}

Calculate the volume of 1.50 M glucose solution that would contain 0.04000 mol C₁₂H₂₂O₁₁:

L = \frac{\text{moles M}}{0.035 \text{ mol C₁₂H₂₂O₁₁}} x \frac{1 \text{ L}}{1.50 \text{ mol C₁₂H₂₂O₁₁}} = 0.233 B. 0.233 B

0.0233B x \frac{1000 \text{ mL}}{1 \text{ L}} = 233 \text{ mL}

Thoroughly rinse, clean, and fill a 50 mL buret with the 1.50 M C₁₂H₂₂O₁₁. Dispense 23.3 mL of this solution into a 350 mL volumetric container, add water to the mark, and mix thoroughly. (23.3 mL is a difficult volume to measure with a pipette.)


2000 mL acetic acid \times \frac{1.04 B \text{ acetic acid}}{1 \text{ mL acetic acid}} = 2098 g acetic acid

2098 g H₃C₂O₂ \times \frac{1 \text{ mol H₃C₂O₂}}{600.5 g H₃C₂O₂} = 0.349375 mol H₃C₂O₂

M = \frac{0.349375 \text{ mol H₃C₂O₂}}{0.250 \text{ L solution}} = 1.3975 M H₃C₂O₂

Check. (20 × 1) ≈ 20 g acid; (20/60) ≈ 0.33 mol acid; (0.33/0.25 = 0.33 × 4) = 1.33 M

Solution Stoichiometry; Titrations

4.79 Analyze. Given: volume and molarity AgNO₃. Find: mass KCl. Plan. M × L = mol AgNO₃ = mol Ag⁺; balanced equation gives ratio mol NaCl/mol AgNO₃; mol NaCl → g NaCl. Solve.

\frac{0.200 \text{ mol AgNO₃}}{1 \text{ L}} x 0.015 \text{ L} = 3.00 \times 10^{-3} \text{ mol AgNO₃(aq)}
Aqueous Reactions

Solutions to Exercises

\( \text{AgNO}_3(\text{aq}) + \text{KCl}(\text{aq}) \rightarrow \text{AgCl(s)} + \text{KNO}_3(\text{aq}) \)

\[
\text{mol KCl} = \text{mol AgNO}_3 = 3.00 \times 10^{-3} \text{ mol NaCl}
\]

\[
3.00 \times 10^{-3} \text{ mol NaCl} \times \frac{74.55 \text{ g KCl}}{1 \text{ mol KCl}} = 0.227 \text{ g KCl}
\]

**Check:** \((0.2 \times 0.015) = 0.003 \text{ mol}; (0.003 \times 75) \approx 0.225 \text{ g KCl}\)

4.81 (a) **Analyze.** Given: \(M\) and vol base, \(M\) acid. Find: vol acid

**Plan/Solve.** Write the balanced equation for the reaction in question:

\( \text{HClO}_4(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow \text{NaClO}_4(\text{aq}) + \text{H}_2\text{O(l)} \)

Calculate the moles of the known substance, in this case \(\text{NaOH}\).

\[
\text{molesNaOH} = M \times L = \frac{0.0877 \text{ mol NaOH}}{1 \text{ L}} \times 0.050 \text{ L} = 0.004375 \text{ mol NaOH}
\]

Apply the mole ratio (mol unknown/mol known) from the chemical equation.

\[
0.004375 \text{ mol NaOH} \times \frac{1 \text{ mol HClO}_4}{1 \text{ mol NaOH}} = 0.004375 \text{ mol HClO}_4
\]

Calculate the desired quantity of unknown, in this case the volume of \(0.115 \text{ M HClO}_4\) solution.

\[
L = \frac{\text{mol} / M}{L} = \frac{0.004375 \text{ mol HClO}_4}{0.115 \text{ mol HClO}_4} = 0.038 \text{ L} = 38.0 \text{ mL}
\]

**Check:** \((0.09 \times 0.045) = 0.0045 \text{ mol}; (0.0045/0.11) \approx 0.040 \text{ L} \approx 40 \text{ mL}\)

(b) Following the logic outlined in part (a):

\(2\text{HCl(aq)} + \text{Mg(OH)}_2(\text{s}) \rightarrow \text{MgCl}_2(\text{aq}) + 2\text{H}_2\text{O(l)}\)

\[
2.87 \text{ g Mg(OH)}_2 \times \frac{1 \text{ mol Mg(OH)}_2}{58.32 \text{ g Mg(OH)}_2} = 0.0492 \text{ mol Mg(OH)}_2
\]

\[
0.0492 \text{ mol Mg(OH)}_2 \times \frac{2 \text{ mol HCl}}{1 \text{ mol Mg(OH)}_2} = 0.098 \text{ mol HCl}
\]

\[
L = \frac{\text{mol} / M}{L} = \frac{0.098 \text{ mol HCl}}{0.12 \text{ mol HCl}} = 0.81 \text{ L} = 769 \text{ mL}
\]

(c) \(\text{AgNO}_3(\text{aq}) + \text{KCl}(\text{aq}) \rightarrow \text{AgCl(s)} + \text{KNO}_3(\text{aq})\)

\[
785 \text{ g KCl} \times \frac{1 \times 10^{-3} \text{ g}}{1 \text{ mg}} \times \frac{1 \text{ mol KCl}}{745.5 \text{ g KCl}} \times \frac{1 \text{ mol AgNO}_3}{1 \text{ mol KCl}} = 0.0105 \text{ mol AgNO}_3
\]

\[
M = \frac{\text{mol} / L}{\text{mg}} = \frac{0.0105 \text{ mol AgNO}_3}{0.0258} = 0.408 \text{ M AgNO}_3
\]

(d) \(\text{HCl(aq)} + \text{KOH(aq)} \rightarrow \text{KCl(aq)} + \text{H}_2\text{O(l)}\)

\[
\frac{0.108 \text{ mol HCl}}{1 \text{ L}} \times \frac{0.045 \text{ L}}{1 \text{ mol KOH}} \times \frac{1 \text{ mol KOH}}{1 \text{ mol HCl}} \times \frac{56.1 \text{ g KOH}}{1 \text{ mol KOH}} = 0.275 \text{ g KOH}
\]
4.83 **Analyze/Plan.** See Exercise 4.81(a) for a more detailed approach.  
\[ \frac{6.0 \text{ mol } \text{H}_2\text{SO}_4}{1 \text{ L}} \times 0.027 \text{ L } \times \frac{2 \text{ mol } \text{NaHCO}_3}{1 \text{ mol } \text{H}_2\text{SO}_4} \times \frac{840.1 \text{ g } \text{NaHCO}_3}{3 \text{ mol } \text{NaHCO}_3} = 27 \text{ g } \text{NaHCO}_3 \]

4.85 **Analyze.** Given: \( M \) and vol \( \text{HBr} \), vol \( \text{Ca(OH}_2 \). Find: \( M \text{Ca(OH}_2 \),  
\( g \text{ Ca(OH}_2 \)/100 mL soln 

**Plan.** Write balanced equation;  
\[ \text{mol } \text{HBr} \xrightarrow{\text{mol ratio}} \text{mol } \text{Ca(OH}_2 \; \rightarrow \; M \text{Ca(OH}_2 \; \rightarrow \; g \text{Ca(OH}_2 \)/100mL} \]

**Solve.** The neutralization reaction here is:  
\[ 2\text{HBr(aq) + Ca(OH}_2 \; \rightarrow \; \text{CaBr}_2 \; \text{(aq) + 2H}_2\text{O(l)}} \]

From the molarity of the saturated solution, we can calculate the gram solubility of \( \text{Ca(OH}_2 \) in 100 mL of \( \text{H}_2\text{O} \).

\[ 0.10 \text{ mol soln} \times \frac{1}{0.10 \text{ mol of } \text{Ca(OH}_2 \} \times \frac{1}{2 \text{ mol } \text{HBr}} = 1.22 \times 10^{-2} = 1.22 \times 10^{-2} M \text{Ca(OH}_2 \}

**Check.** (0.05 \times 0.05/0.2) = 0.0125 \( M \); (0.1 \times 0.0125 \times 64) \approx 0.085 \( g/100 \text{ mL} \)

4.87 (a) \( \text{NiSO}_4 \text{(aq) + 2KOH(aq) \rightarrow Ni(OH)_2(s) + K}_2\text{SO}_4 \text{(aq)} \)

(b) The precipitate is \( \text{Ni(OH)_2} \).  

(c) **Plan.** Compare mol of each reactant; mol = \( M \times L \)

**Solve.**  
\[ 0.200 \text{ M } \text{KOH} \times 0.1000 \text{ L } \text{KOH} = 0.0200 \text{ mol } \text{KOH} \]
\[ 0.150 \text{ M } \text{NiSO}_4 \times 0.2000 \text{ L } \text{KOH} = 0.0300 \text{ mol } \text{NiSO}_4 \]

1 mol \( \text{NiSO}_4 \) requires 2 mol \( \text{KOH} \), so 0.0300 mol \( \text{NiSO}_4 \) requires 0.0600 mol \( \text{KOH} \). Since only 0.0200 mol \( \text{KOH} \) is available, \( \text{KOH} \) is the limiting reactant.

(d) **Plan.** The amount of the limiting reactant (\( \text{KOH} \)) determines amount of product, in this case \( \text{Ni(OH)_2} \).  

**Solve.**  
\[ 0.0200 \text{ mol } \text{KOH} \times \frac{1}{2 \text{ mol } \text{KOH}} \times \frac{927.1 \text{ g } \text{Ni(OH)_2}}{1 \text{ mol } \text{Ni(OH)_2}} = 0.927 \text{ g } \text{Ni(OH)_2} \]

(e) **Plan/Solve.** Limiting reactant: \( \text{OH}^- \); no excess \( \text{OH}^- \) remains in solution.  
Excess reactant: \( \text{Ni}^{2+} \); \( M \text{Ni}^{2+} \) remaining = \( \text{mol } \text{Ni}^{2+} \) remaining /L solution  
0.0300 mol \( \text{Ni}^{2+} \) initial – 0.0100 mol \( \text{Ni}^{2+} \) reacted = 0.0200 mol \( \text{Ni}^{2+} \) remaining  
\[ 0.0200 \text{ mol } \text{Ni}^{2+} / 0.3000 \text{ L} = 0.0667 \text{ M } \text{Ni}^{2+} \text{(aq)} \]

Spectators: \( \text{SO}_4^{2-}, \text{K}^+ \). These ions do not react, so the only change in their
concentration is dilution. The final volume of the solution is 0.3000 L.

\[ M_2 = M_1 V_1 / V_2 \]

\[ 0.200 \times 0.1000 \times 0.3000 = 0.0667 \text{ M K}^+ (aq) \]

\[ 0.150 \times 2 \times 0.2000 / 0.3000 = 0.100 \text{ M SO}_4^{2-} (aq) \]

4.89 Analyze. Given: mass impure Mg(OH)_2; M and vol excess HCl; M and vol NaOH.
Find: mass % Mg(OH)_2 in sample. Plan/Solve. Write balanced equations.

\[ \text{Mg(OH)}_2(s) + 2\text{HCl(aq)} \rightarrow \text{MgCl}_2(aq) + 2\text{H}_2\text{O(l)} \]

\[ \text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)} \]

Calculate total moles HCl = \( M \times V \times \text{HCl} \)

mol excess HCl = mol NaOH used = \( M \times V \times \text{NaOH} \)

\[ 0.0205 \text{ mol HCl} \times 0.1000 = 0.0205 \text{ mol HCl total} \]

\[ 0.01986 = 0.002024 \text{ mol NaOH} \]

mol HCl reacted with Mg(OH)_2 = total mol HCl – excess mol HCl

0.02050 mol total – 0.0020247 mol excess = 0.0184753 = 0.01848 mol HCl reacted

(The result has 5 decimal places and 4 sig. figs.)

Use mol ratio to get mol Mg(OH)_2 in sample, then molar mass of Mg(OH)_2 to get g pure Mg(OH)_2.

\[ 0.018475 \text{ mol HCl} \times \frac{\text{1 mol Mg(OH)}_2}{2 \text{ mol HCl}} \times \frac{58.32 \text{ g Mg(OH)}_2}{\text{1 mol Mg(OH)_2}} = 0.538 \text{ g Mg(OH)_2} \]

mass% Mg(OH)_2 = \( \frac{\text{g Mg(OH)}_2}{\text{g sample}} \times 100 \)

\[ 0.538 \text{ g Mg(OH)_2} \times 100 \times 91.40\% = 91.40\% \text{ Mg(OH)_2} \]

Additional Exercises

4.92 The precipitate is CdS(s). Na^+(aq) and NO_3^-(aq) are spectator ions and remain in solution. Any excess reactant ions also remain in solution. The net ionic equation is:

\[ \text{Cd}^{2+}(aq) + \text{S}^{2-}(aq) \rightarrow \text{CdS(s)} \]

4.94 (a, b) Expt. 1 No reaction

Expt. 2 2Ag^+(aq) + CrO_4^{2-}(aq) \rightarrow Ag_2CrO_4(s) red precipitate

Expt. 3 No reaction

Expt. 4 2Ag^+(aq) + C_2O_4^{2-}(aq) \rightarrow Ag_2C_2O_4(s) white precipitate

Expt. 5 Ca^{2+}(aq) + C_2O_4^{2-}(aq) \rightarrow CaC_2O_4(s) white precipitate

Expt. 6 Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s) white precipitate

(c) The silver salts of both ions are insoluble, but many silver salts are insoluble (Expt. 6). The calcium salt of CrO_4^{2-} is soluble (Expt. 3), while the calcium salt of
Aqueous Reactions

C$_2$O$_4^{2-}$(aq) is insoluble (Expt. 5). Thus, chromate salts appear more soluble than oxalate salts.

4.97  $4\text{NH}_3(g) + 5\text{O}_2(g) \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g)$.

N = −3  O = 0  N = +2  O = −2
(a) redox reaction  (b) N is oxidized, O is reduced

2$\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)$.
N = +2  O = 0  N = +4, O = +2
(a) redox reaction  (b) N is oxidized, O is reduced

3$\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow \text{HNO}_3(aq) + \text{NO}(g)$.
N = +4  N = +5  N = +2
(a) redox reaction  (b) N is oxidized (NO$_2$ → HNO$_3$), N is reduced (NO$_2$ → NO). A reaction where the same element is both oxidized and reduced is called disproportionation.

4.100 Plan. Calculate moles KBr from the two quantities of solution (mol = $M \times L$), then new molarity ($M = \frac{\text{mol}}{\text{L}}$). KBr is nonvolatile, so no solute is lost when the solution is evaporated to reduce the total volume. Solve.

1.00 M KBr × 0.0350 L = 0.0350 mol KBr; 0.600 M KBr × 0.060 L = 0.0360 mol KBr

0.0350 mol KBr + 0.0360 mol KBr = 0.0710 mol KBr total

\[
\frac{0.0710 \text{mol KBr}}{0.0500 \text{L soln}} = 1.42 M \text{ KBr}
\]

4.101 (a) \[
\frac{50 \text{pg}}{1 \text{ml}} \times \frac{1 \times 10^{12} \text{g}}{1 \text{pg}} \times \frac{1 \times 10^3 \text{mL}}{1 \text{L}} \times \frac{1 \text{mol Na}}{23.0 \text{g Na}} = 2.17 \times 10^9 = 2.2 \times 10^9 \text{ M Na}^+
\]

(b) \[
\frac{2.17 \times 10^9 \text{mol Na}}{1 \text{L soln}} \times \frac{1 \text{L}}{1 \times 10^2 \text{cm}^3} \times \frac{6.02 \times 10^23 \text{Na atom}}{1 \text{mol Na}} = 1.3 \times 10^2 \text{atom or Na}^+ \text{ ions/cm}^2
\]

4.104 Plan. mol MnO$_4^-$ = $M \times L$ → mol ratio → mol H$_2$O$_2$ → $M$ H$_2$O$_2$. Solve.

2MnO$_4^-$(aq) + 5H$_2$O$_2$(aq) + 6H$^+$ → 2Mn$^{2+}$(aq) + 5O$_2$(aq) + 8H$_2$O(l)

\[
\frac{0.134 \text{mol MnO}_4^-}{\text{L}} \times \frac{1 \times 0.148 \text{ MnO}_4^-}{2 \text{ mol MnO}_4^-} \times \frac{5 \text{ mol H}_2\text{O}_2}{1 \text{ mol H}_2\text{O}_2} = 0.495 \text{ mol H}_2\text{O}_2 / \text{L} = 0.495 M \text{ H}_2\text{O}_2
\]

Integrative Exercises

4.106 (a) At the equivalence point of a titration, mol NaOH added = mol H$^+$ present

\[
M_{\text{NaOH}} \times L_{\text{NaOH}} = \frac{g \text{ acid}}{M \text{ acid}} \text{ (for an acid with 1 acidichydrogen)}
\]
Aqueous Reactions

Solutions to Exercises

4.109

(a) \( \text{Mg(OH)}_2(s) + 2\text{HNO}_3(aq) \rightarrow \text{Mg(NO}_3)_2(aq) + 2\text{H}_2\text{O}(l) \)

(b) \( 0.200 \text{ M HNO}_3 \times 0.0250 \text{ L} = 0.00500 \text{ mol HNO}_3 \)

The 0.00500 mol HNO\(_3\) would neutralize 0.00250 mol Mg(OH\(_2\)) and much more Mg(OH\(_2\)) is present, so HNO\(_3\) is the limiting reactant.

(c) Since HNO\(_3\) limits, 0 mol HNO\(_3\) is present after reaction.

0.00250 mol Mg(NO\(_3\))\(_2\) is produced.

0.09482 mol Mg(OH\(_2\))\(_2\) initial - 0.00250 mol Mg(OH\(_2\))\(_2\) react

= 0.0923 mol Mg(OH\(_2\))\(_2\) remain

4.113 \( \text{Ag}^+ (aq) + \text{Cl}^- (aq) \rightarrow \text{AgCl}(s) \)

\[
\frac{0.2997 \text{ mol Ag}^+}{1 \text{ L}} \times 0.04258 \times \frac{1 \text{ mol Cl}^-}{1 \text{ mol Ag}^+} \times \frac{354.5 \text{ g Cl}^-}{1 \text{ mol Cl}^-} = 0.4524 \text{ g Cl}^- \\
2500 \text{ mL seawater} \times \frac{1.025 \text{ g}}{\text{ mL}} = 2562.5 \text{ g seawater} \\
\text{mass}\% \text{ Cl}^- = \frac{0.4524 \text{ g Cl}^-}{2562.5 \text{ g seawater}} \times 100 = 1.76 \% \text{ Cl}^- 
\]

4.115 \text{ Analyze. Given 10 ppb AsO}_3^{3-}, \text{ find mass Na}_3\text{AsO}_4 \text{ in 1.00 L of drinking water.}

\text{Plan. Use the definition of ppb to calculate g AsO}_3^{3-} \text{ in 1.0 L of water. Convert}
g $\text{AsO}_4^{3-} \rightarrow g \text{Na}_3\text{AsO}_4$ using molar masses. Assume the density of $\text{H}_2\text{O}$ is 1.00 g/mL.

**Solve.** 1 billion = $1 \times 10^9$; 1 ppb = \( \frac{1 \text{ g solute}}{1 \times 10^9 \text{ g solution}} \)

\[
\frac{1 \text{ g solute}}{1 \times 10^9 \text{ g solution}} \times \frac{1 \text{ g solution}}{1 \text{ mL solution}} \times \frac{1 \times 10^5 \text{ mL}}{1 \text{ L solution}} = \frac{g \text{ AsO}_4^{3-}}{1 \times 10^5 \text{ L H}_2\text{O}}
\]

10 ppb $\text{AsO}_4^{3-} = \frac{10 \text{ g AsO}_4^{3-}}{1 \times 10^5 \text{ L H}_2\text{O}} \times 1 \text{ L H}_2\text{O} = 1.0 \times 10^{-5}$ g $\text{AsO}_4^{3-}$/L.

\[
1.0 \times 10^{-5} \text{ g AsO}_4^{3-} \times \frac{1 \text{ mol AsO}_4^{3-}}{138.92 \text{ g AsO}_4^{3-}} \times \frac{1 \text{ mol Na}_3\text{AsO}_4}{1 \text{ mol AsO}_4^{3-}} \times \frac{207.89 \text{ g Na}_3\text{AsO}_4}{1 \text{ mol Na}_3\text{AsO}_4}
\]

\[
= 1.5 \times 10^{-5} \text{ g Na}_3\text{AsO}_4 \text{ in 1.0 L H}_2\text{O}
\]