

**Supplemental Questions
for
Reactions in Aqueous Solution**

1a. What is the $[\text{NH}_4^+]$ in a solution that was prepared by dissolving 7.43 g $(\text{NH}_4)_2\text{SO}_4$ in 250.0 mL of solution? The molar mass of $(\text{NH}_4)_2\text{SO}_4$ is 132.13 g/mole.

$$7.43 \text{ g } (\text{NH}_4)_2\text{SO}_4 \left(\frac{1 \text{ mole } (\text{NH}_4)_2\text{SO}_4}{132.13 \text{ g } (\text{NH}_4)_2\text{SO}_4} \right) \left(\frac{2 \text{ mole } \text{NH}_4^+}{1 \text{ mole } (\text{NH}_4)_2\text{SO}_4} \right) = 0.112_{46} \text{ mole } \text{NH}_4^+$$

$$[\text{NH}_4^+] = \frac{0.112_{46} \text{ mole } \text{NH}_4^+}{250.0 \times 10^{-3} \text{ L}} = 0.450 \text{ M}$$

The $[\text{NH}_4^+]$ is 0.450 M.

b. What quantity of potassium phosphate, in grams, is required to prepare 500.0 mL of a solution where $[\text{K}^+]$ is 0.100 M? The molar mass of potassium phosphate is 212.28 g/mole.

$$500.0 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{0.100 \text{ mole } \text{K}^+}{1 \text{ L}} \right) \left(\frac{1 \text{ mole } \text{K}_3\text{PO}_4}{3 \text{ mole } \text{K}^+} \right) \left(\frac{212.28 \text{ g } \text{K}_3\text{PO}_4}{1 \text{ mole } \text{K}_3\text{PO}_4} \right) = 3.54 \text{ g } \text{K}_3\text{PO}_4$$

To prepare 500.0 mL of a solution with $[\text{K}^+] = 0.100 \text{ M}$, one needs 3.54 g K_3PO_4 .

c. Ringer's solution, used in the treatment of burns and wounds, is prepared by dissolving 4.30 g NaCl, 0.150 g KCl and 0.165 g CaCl_2 in water and diluting to 500.0 mL. What is the molarity of each ion present in solution? Molar masses: NaCl 58.44 g/mole, KCl 74.56 g/mole, and CaCl_2 110.99 g/mole.

Calculate moles of each salt present in the solution.

$$4.30 \text{ g NaCl} \left(\frac{1 \text{ mole}}{57.44 \text{ g}} \right) = 0.735_8 \text{ mole NaCl}$$

$$0.150 \text{ g KCl} \left(\frac{1 \text{ mole}}{74.56 \text{ g}} \right) = 0.00201_2 \text{ mole KCl}$$

$$0.165 \text{ g CaCl}_2 \left(\frac{1 \text{ mole}}{110.99 \text{ g}} \right) = 0.00148_7 \text{ mole CaCl}_2$$

From this we can calculate the concentrations of each ion in solution.

$$[\text{Na}^+] = \frac{0.735_8 \text{ mole Na}^+}{0.5000 \text{ L}} = 0.147 \text{ M}$$

$$[\text{K}^+] = \frac{0.00201_2 \text{ mole K}^+}{0.5000 \text{ L}} = 4.02 \times 10^{-3} \text{ M}$$

$$[Ca^{2+}] = \frac{0.001487 \text{ mole } Ca^{2+}}{0.5000 \text{ L}} = 2.97 \times 10^{-3} \text{ M}$$

$$[Cl^-] = \frac{0.07358 + 0.002012 + 2(0.001487) \text{ mole } Cl^-}{0.5000 \text{ L}} = 0.157 \text{ M}$$

In Ringer's solution the $[Na^+]$ is 0.147 M, the $[K^+]$ is 4.02×10^{-3} M, the $[Ca^{2+}]$ is 2.07×10^{-3} M and the $[Cl^-]$ is 0.157 M.

d. Trandate[®] is a drug used to treat hypertension. It is sold as a solution containing 5.00 mg of the drug per 1.00 mL of solution. What volume, in mL, of this solution do you need to prepare 500.0 mL of a new solution containing 0.0100 mg of the drug per 1.00 mL of solution?

This is a simple dilution problem. So, use equation $C_i V_i = C_d V_d$

$$C_i V_i = C_d V_d$$

$$\left(\frac{5.00 \text{ mg}}{1.00 \text{ mL}} \right) V_i = \left(\frac{0.0100 \text{ mg}}{1.00 \text{ mL}} \right) (500.0 \text{ mL})$$

$$V_i = 1.00 \text{ mL}$$

One needs 1.00 mL of the original solution to prepare 500.0 mL of a solution with a concentration of 0.0100 mg/mL.

e. You need 500.0 mL of a solution where $[Ni^{2+}]$ is 1.00×10^{-6} M, and you have a solution where the $[Ni^{2+}]$ is 0.100 M. At your disposal are a 500.0-mL and a 1000.0-mL volumetric flask, and a 1.000-mL, a 2.000-mL, a 5.000-mL and a 10.00-mL pipette. What serial dilution do you need to perform to prepare the more dilute solution from the more concentrated one?

This is a twist on a standard serial dilution problem, and is one that is often encountered in the laboratory. We need to prepare a solution by serial dilution starting from a stock solution, but we need to find the volumes to take. As always, start with the serial dilution equation.

$$\left(\frac{V_1}{1000.0 \text{ mL}} \right) \left(\frac{V_2}{500.0 \text{ mL}} \right) (0.100 \text{ M}) = 1.00 \times 10^{-6} \text{ M}$$

$$V_1 V_2 = 5.00 \text{ mL}^2$$

With the pipettes given $V_1 = 1.000$ mL and $V_2 = 5.000$ mL or vice versa.

The dilutions that need to be done are thus, 1.000 mL of the 0.100 M solution is diluted to 1000.0 mL and then 5.000 mL of this solution is diluted to 500.0 mL.

f. A solution is prepared by dissolving 15.347 g Na_3PO_4 and 2.771 g Na_2HPO_4 in water to make 500.0 mL of solution. What is $[\text{Na}^+]$ in the solution? The formula weight of Na_3PO_4 is 163.94 g/mole, and the formula weight of Na_2HPO_4 is 119.68 g/mole.

The Na^+ comes from two sources, the Na_3PO_4 and the Na_2HPO_4 , so the total amount of Na^+ in solution is the sum of the moles of Na^+ from each source. The concentration, in molar, is simply the moles of Na^+ divided by the volume of the solution in liters.

$$15.347 \text{ g } \text{Na}_3\text{PO}_4 \left(\frac{1 \text{ mole } \text{Na}_3\text{PO}_4}{163.94 \text{ g } \text{Na}_3\text{PO}_4} \right) \left(\frac{3 \text{ mole } \text{Na}^+}{1 \text{ mole } \text{Na}_3\text{PO}_4} \right) = 0.28084_1 \text{ mole } \text{Na}^+$$

$$2.771 \text{ g } \text{Na}_2\text{HPO}_4 \left(\frac{1 \text{ mole } \text{Na}_2\text{HPO}_4}{119.68 \text{ g } \text{Na}_2\text{HPO}_4} \right) \left(\frac{2 \text{ mole } \text{Na}^+}{1 \text{ mole } \text{Na}_2\text{HPO}_4} \right) = 0.04630_7 \text{ mole } \text{Na}^+$$

$$\text{Total number of moles } \text{Na}^+ = 0.28084_1 + 0.04630_7 = 0.32714_8 \text{ mole}$$

$$[\text{Na}^+] = \left(\frac{0.32714_8 \text{ mole } \text{Na}^+}{500.0 \times 10^{-3} \text{ L}} \right) = 0.6543 \text{ M } \text{Na}^+$$

The concentration of Na^+ in the solution is 0.6543 M.

g. An aqueous methanol solution is 10.0% by volume methanol. What is the molarity of the solution? The density of methanol is 0.791 g/mL and the molar mass of methanol is 32.04 g/mole.

Use the definition of % by volume to rewrite the given percent in terms of mL of methanol per 100.0 mL of solution.

$$\left(\frac{10.0 \text{ mL methanol}}{100.0 \times 10^{-3} \text{ L solution}} \right) \left(\frac{0.791 \text{ g methanol}}{1 \text{ mL methanol}} \right) \left(\frac{1 \text{ mole methanol}}{32.04 \text{ g methanol}} \right) = 2.47 \text{ M methanol}$$

The molarity of methanol of a 10.0% by volume solution is 2.47 M.

h. You have a 0.250 M $\text{Ca}(\text{NO}_3)_2$ solution, and you need 500.0 mL of a solution where $[\text{NO}_3^-]$ is 0.0125 M. What dilution do you perform?

Key it to remember that for every 1 mole of $\text{Ca}(\text{NO}_3)_2$ there are 2 moles of NO_3^- .

$$C_i V_i = C_d V_d$$

$$\left(\frac{0.250 \text{ mole } \text{Ca}(\text{NO}_3)_2}{1 \text{ L}} \right) \left(\frac{2 \text{ mole } \text{NO}_3^-}{1 \text{ mole } \text{Ca}(\text{NO}_3)_2} \right) V_i = (500.0 \text{ mL}) (0.0125 \text{ M } \text{NO}_3^-)$$

$$V_i = 12.5 \text{ mL}$$

Dilute 12.5 mL of the $\text{Ca}(\text{NO}_3)_2$ solution to 500.0 mL with water.

i. You need to prepare 250.0 mL of a solution with a $[Fe^{3+}]$ of 0.0125 M. How many grams of $Fe(NO_3)_3 \cdot 9H_2O$ do you need? The molar mass of $Fe(NO_3)_3 \cdot 9H_2O$ is 404.00 g/mole and the molar mass of H_2O is 18.015 g/mole.

This problem is a straight-forward solution preparation from a solid reagent.

$$250.0 \times 10^{-3} L \left(\frac{0.0125 \text{ mole } Fe^{3+}}{1 L} \right) \left(\frac{1 \text{ mole } Fe(NO_3)_3 \cdot 9 H_2O}{1 \text{ mole } Fe^{3+}} \right) \left(\frac{404.00 \text{ g } Fe(NO_3)_3 \cdot 9 H_2O}{1 \text{ mole } Fe(NO_3)_3 \cdot 9 H_2O} \right) = 1.26 \text{ g } Fe(NO_3)_3 \cdot 9 H_2O$$

To prepare the given solution, 1.26 g $Fe(NO_3)_3 \cdot 9H_2O$ must be dissolved in water and diluted to 250.0 mL.

j. How many grams of $(NH_4)_2SO_4$ do you need to prepare 250.0 mL of a solution where $[NH_4^+] = 0.450 M$? The molar mass of $(NH_4)_2SO_4$ is 132.13 g/mole.

$$250.0 \text{ mL} \left(\frac{1 L}{1000 \text{ mL}} \right) \left(\frac{0.450 \text{ mole } NH_4^+}{1 L} \right) \left(\frac{1 \text{ mole } (NH_4)_2SO_4}{2 \text{ mole } NH_4^+} \right) \left(\frac{132.13 \text{ g}}{1 \text{ mole } (NH_4)_2SO_4} \right) = 7.43 \text{ g}$$

To prepare this solution 7.43 g of $(NH_4)_2SO_4$ are needed.

k. A 50.00 mL portion of a cresol red solution with a concentration of $4.000 \times 10^{-2} M$ is diluted to 100.0 mL. A 25.00 mL portion of this new solution is then diluted to 250.0 mL to give a third solution. What is the concentration cresol red in the third solution?

$$\left(\frac{50.00 \text{ mL}}{100.0 \text{ mL}} \right) \left(\frac{25.00 \text{ mL}}{250.0 \text{ mL}} \right) (4.000 \times 10^{-2} M) = 2.000 \times 10^{-3} M$$

The concentration of cresol red in the third solution is $2.000 \times 10^{-3} M$.

l. How many mL of a 0.500 M NaCl solution are needed to prepare 500.0 mL of a solution where $[NaCl]$ is 0.0250 M?

$$\begin{aligned} C_i \cdot V_i &= C_d \cdot V_d \\ C_i (0.500 M) &= (500.0 \text{ mL})(0.0250 M) \\ C_i &= 25.0 \text{ mL} \end{aligned}$$

To prepare this solution one needs 25.0 mL of the 0.500 M NaCl.

m. A 10.00 mL portion of a 0.500 M sodium sulfate solution is diluted to 500.0 mL. What is the $[\text{Na}^+]$ ion in the new solution?

Find moles of Na^+ taken from the 0.500 M solution, and then use the definition of molarity to calculate the new concentration.

$$10.00 \times 10^{-3} \text{ L } \text{Na}_2\text{SO}_4 \text{ solution} \left(\frac{0.500 \text{ mole Na}_2\text{SO}_4}{1 \text{ L Na}_2\text{SO}_4 \text{ solution}} \right) \left(\frac{2 \text{ mole Na}^+}{1 \text{ mole Na}_2\text{SO}_4} \right) = 0.0100 \text{ mole Na}^+$$

$$[\text{Na}^+] = \frac{0.0100 \text{ mole Na}^+}{500.0 \times 10^{-3} \text{ L}} = 0.0200 \text{ M}$$

The $[\text{Na}^+]$ in the new solution is 0.0200 M.

Alternate solution. Solved using the dilution equation, where the quantity in the square brackets is the $[\text{Na}^+]$ in the original solution.

$$C_i \cdot V_i = C_d \cdot V_d$$

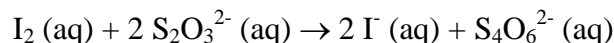
$$\left[\left(\frac{0.500 \text{ mole Na}_2\text{SO}_4}{1 \text{ L solution}} \right) \left(\frac{2 \text{ mole Na}^+}{1 \text{ mole Na}_2\text{SO}_4} \right) \right] (10.0 \text{ mL}) = C_d (500.0 \text{ mL})$$

$$C_d = \frac{\left[\left(\frac{0.500 \text{ mole Na}_2\text{SO}_4}{1 \text{ L solution}} \right) \left(\frac{2 \text{ mole Na}^+}{1 \text{ mole Na}_2\text{SO}_4} \right) \right] (10.0 \text{ mL})}{500.0 \text{ mL}} = 0.0200 \text{ M Na}^+$$

Another alternate solution. This one is solved like a serial dilution problem, where what's in the square brackets gives the $[\text{Na}^+]$ in the original solution.

$$\frac{10.0 \text{ mL}}{500.0 \text{ mL}} \left[\left(\frac{0.500 \text{ mole Na}_2\text{SO}_4}{1 \text{ L solution}} \right) \left(\frac{2 \text{ mole Na}^+}{1 \text{ mole Na}_2\text{SO}_4} \right) \right] = 0.0200 \text{ M Na}^+$$

2. Sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$, is used as a “fixer” in photography. A friend who is a photographer gives you an old bottle of fixer solution and wants you to determine its purity. After consulting a book on quantitative analysis you find that the thiosulfate ion can be oxidized with I_2 according to the following balanced net ionic equation.



You titrate 3.232 g of the fixer solution with a 0.246 M I_2 solution and find that 40.21 mL of the I_2 solution is required to reach the equivalence point. What is the weight percent of $\text{Na}_2\text{S}_2\text{O}_3$ in the fixer solution?

	I_2	$2 \text{S}_2\text{O}_3^{2-}$
Volume (mL)	40.21	
Concentration (M)	0.246	
Mass (g)	-----	158.110 (for $\text{Na}_2\text{S}_2\text{O}_3$)
Molar mass (g/mole)	-----	
Moles	$9.89_2 \times 10^{-6}$	$1.97_8 \times 10^{-2}$

Calculate moles of I_2 in the given volume.

$$40.21 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{0.246 \text{ mole } \text{I}_2}{1 \text{ L}} \right) = 9.89_2 \times 10^{-3} \text{ mole } \text{I}_2$$

This is the amount of I_2 to reach the equivalence point, which means there is exactly the number of moles of I_2 needed to react with the $\text{S}_2\text{O}_3^{2-}$ present in the fixer.

$$9.89_2 \times 10^{-3} \text{ mole } \text{I}_2 \left(\frac{2 \text{ S}_2\text{O}_3^{2-}}{1 \text{ mole } \text{I}_2} \right) = 1.97_8 \times 10^{-2} \text{ mole } \text{S}_2\text{O}_3^{2-}$$

Now convert moles of $\text{S}_2\text{O}_3^{2-}$ to grams of $\text{Na}_2\text{S}_2\text{O}_3$.

$$1.97_8 \times 10^{-2} \text{ mole } \text{S}_2\text{O}_3^{2-} \left(\frac{1 \text{ mole } \text{Na}_2\text{S}_2\text{O}_3}{1 \text{ mole } \text{S}_2\text{O}_3^{2-}} \right) \left(\frac{158.11 \text{ g } \text{Na}_2\text{S}_2\text{O}_3}{1 \text{ mole } \text{Na}_2\text{S}_2\text{O}_3} \right) = 3.13 \text{ g } \text{Na}_2\text{S}_2\text{O}_3$$

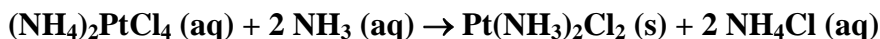
Calculate % $\text{Na}_2\text{S}_2\text{O}_3$ by weight in the fixer.

$$\left(\frac{3.13 \text{ g } \text{Na}_2\text{S}_2\text{O}_3}{3.232 \text{ g fixer}} \right) \times 100 = 96.8\%$$

The fixer contains 96.8% $\text{Na}_2\text{S}_2\text{O}_3$ by weight.

3. The cancer chemotherapy drug cisplatin, $\text{Pt}(\text{NH}_3)_2\text{Cl}_2$, can be made by reacting an aqueous solution of $(\text{NH}_4)_2\text{PtCl}_4$ with an aqueous solution of NH_3 . The other product of the reaction is ammonium chloride.

a. Write the balanced chemical reaction for this reaction, given that $\text{Pt}(\text{NH}_3)_2\text{Cl}_2$ is not very soluble in water.



b. Platinum is more expensive than gold, so it is imperative when working with platinum that it all reacts. If we want to obtain a theoretical yield of 12.50 g $\text{Pt}(\text{NH}_3)_2\text{Cl}_2$, how many grams of $(\text{NH}_4)_2\text{PtCl}_4$ and how many mL of a 0.125 M NH_3 solution do we need? Molar masses: ammonium chloride 53.49 g/mole, $(\text{NH}_4)_2\text{PtCl}_4$ 372.96 g/mole and $\text{Pt}(\text{NH}_3)_2\text{Cl}_2$ 300.04 g/mole.

	$(\text{NH}_4)_2\text{PtCl}_4$	2 NH_3	$\text{Pt}(\text{NH}_3)_2\text{Cl}_2$	NH_4Cl
Volume (mL)	-----	?	-----	-----
Concentration (M)	-----	0.125	-----	-----
Mass (g)	?	-----	12.50	-----
Molar mass (g/mole)	372.96	-----	300.04	-----
Moles				-----

Calculate the number of moles of $\text{Pt}(\text{NH}_3)_2\text{Cl}_2$.

$$12.50 \text{ g } \text{Pt}(\text{NH}_3)_2\text{Cl}_2 \left(\frac{1 \text{ mole } \text{Pt}(\text{NH}_3)_2\text{Cl}_2}{300.04 \text{ g } \text{Pt}(\text{NH}_3)_2\text{Cl}_2} \right) = 4.166_1 \times 10^{-2} \text{ mole } \text{Pt}(\text{NH}_3)_2\text{Cl}_2$$

Calculate moles of NH_3 , and then the volume of the solution.

$$4.166 \times 10^{-2} \text{ mole } \text{Pt}(\text{NH}_3)_2\text{Cl}_2 \left(\frac{2 \text{ mole } \text{NH}_3}{1 \text{ mole } \text{Pt}(\text{NH}_3)_2\text{Cl}_2} \right) \left(\frac{1 \text{ L}}{0.125 \text{ mole } \text{NH}_3} \right) \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right) =$$

$$= 667. \text{ mL of the } \text{NH}_3 \text{ solution}$$

Calculate moles, and then the number of grams, of $(\text{NH}_4)_2\text{PtCl}_4$.

$$4.166_1 \times 10^{-2} \text{ mole } \text{Pt}(\text{NH}_3)_2\text{Cl}_2 \left(\frac{1 \text{ mole } (\text{NH}_4)_2\text{PtCl}_4}{1 \text{ mole } \text{Pt}(\text{NH}_3)_2\text{Cl}_2} \right) \left(\frac{372.96 \text{ g } (\text{NH}_4)_2\text{PtCl}_4}{1 \text{ mole } (\text{NH}_4)_2\text{PtCl}_4} \right) =$$

$$= 15.54 \text{ g } (\text{NH}_4)_2\text{PtCl}_4$$

To prepare 12.50 g of cisplatin 15.54 g $(\text{NH}_4)_2\text{PtCl}_4$ and 667. mL of a 0.125 M NH_3 solution are needed.

4. Titanium(IV) chloride reacts with water according to the following balanced equation:
 $\text{TiCl}_4 (\text{l}) + 2 \text{H}_2\text{O} (\text{l}) \rightarrow \text{TiO}_2 (\text{s}) + 4 \text{HCl} (\text{g})$

a. If 14.0 g of TiCl_4 and 14.0 g of H_2O are allowed to react, what is the maximum amount of TiO_2 , in grams, that may be formed?

Molar masses: TiCl_4 189.71 g/mole, H_2O 18.02 g/mole, and TiO_2 79.90 g/mole.

Find moles of each reactant and predict mass of TiO_2 to determine limiting reagent.

$$14.0 \text{gTiCl}_4 \left(\frac{1 \text{moleTiCl}_4}{189.71 \text{g}} \right) \left(\frac{1 \text{moleTiO}_2}{1 \text{moleTiCl}_4} \right) \left(\frac{79.90 \text{gTiO}_2}{1 \text{moleTiO}_2} \right) = 5.90 \text{gTiO}_2$$

$$14.0 \text{gH}_2\text{O} \left(\frac{1 \text{moleH}_2\text{O}}{18.02 \text{gH}_2\text{O}} \right) \left(\frac{1 \text{moleTiO}_2}{2 \text{moleH}_2\text{O}} \right) \left(\frac{79.90 \text{gTiO}_2}{1 \text{moleTiO}_2} \right) = 31.0 \text{gTiO}_2$$

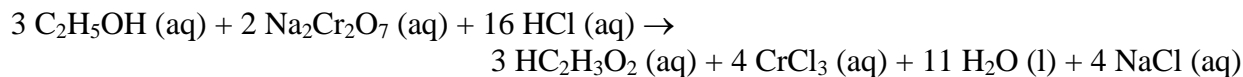
The TiCl_4 is the limiting reagent and the maximum amount of TiO_2 that can be formed in this reaction is 5.90 g.

b. If 4.79 g of TiO_2 were collected from the reaction, what was the percent yield for the reaction?

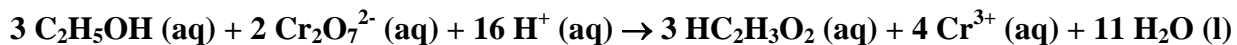
$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{4.79 \text{ g}}{5.90 \text{ g}} \times 100 = 81\%$$

The % yield of this reaction is 81%.

5. Ethanol reacts with sodium dichromate in an aqueous hydrochloric acid solution according to the balanced chemical equation shown below.



a. What is the net ionic equation for this reaction?



b. What kind of reaction is this? **It is a redox reaction.**

c. If 1.00 mL of C_2H_5OH (density 0.816 g/mL) is allowed to react with 32.76 mL of a solution that is 0.5775 M in $Na_2Cr_2O_7$ and 6.000 M in HCl , how much $CrCl_3$, in grams, will be formed? The molar mass of ethanol is 46.07 g/mole, and the molar mass of $CrCl_3$ is 158.36 g/mole.

Find moles of each reactant and determine limiting reagent.

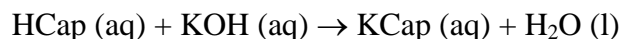
$$1.00 \text{ mL } C_2H_5OH \left(\frac{0.816 \text{ g}}{1 \text{ mL}} \right) \left(\frac{1 \text{ mole } C_2H_5OH}{46.07 \text{ g } C_2H_5OH} \right) \left(\frac{4 \text{ mole } CrCl_3}{3 \text{ mole } C_2H_5OH} \right) \left(\frac{158.36 \text{ g } CrCl_3}{1 \text{ mole } CrCl_3} \right) = 3.74 \text{ g } CrCl_3$$

$$32.76 \times 10^{-3} \text{ L} \left(\frac{0.5775 \text{ mole } Na_2Cr_2O_7}{1 \text{ L}} \right) \left(\frac{4 \text{ mole } CrCl_3}{2 \text{ mole } Na_2Cr_2O_7} \right) \left(\frac{158.36 \text{ g } CrCl_3}{1 \text{ mole } CrCl_3} \right) = 5.992 \text{ g } CrCl_3$$

$$32.76 \times 10^{-3} \text{ L} \left(\frac{6.000 \text{ mole } HCl}{1 \text{ L}} \right) \left(\frac{4 \text{ mole } CrCl_3}{16 \text{ mole } HCl} \right) \left(\frac{158.36 \text{ g } CrCl_3}{1 \text{ mole } CrCl_3} \right) = 7.782 \text{ g } CrCl_3$$

C_2H_5OH is the limiting reagent and 3.74 g $CrCl_3$ will be produced in the reaction.

6a. Caproic acid (abbreviated HCap) is a monoprotic weak acid that smells like wet goat. Titrating 0.2334 g of caproic acid with 0.1004 M KOH requires 20.012 mL of the KOH solution to reach the equivalence point (reaction given below). What is the molecular weight of caproic acid?



	HCap (aq)	KOH (aq)	KCap (aq)	H ₂ O (l)
Mass (g)	0.2334	-----	-----	-----
Molar mass (g/mole)	?	-----	-----	-----
Concentration (M)	-----	0.1004	-----	-----
Volume (mL)	-----	20.012	-----	-----
Moles	$2.009_2 \times 10^{-3}$	$2.009_2 \times 10^{-3}$	-----	-----

$$20.012 \times 10^{-3} \text{ L } KOH \left(\frac{0.1004 \text{ mole } KOH}{1 \text{ L}} \right) \left(\frac{1 \text{ mole } HCap}{1 \text{ mole } KOH} \right) = 2.009_2 \times 10^{-3} \text{ mole } HCap$$

$$\text{Molar mass} = \frac{0.2334 \text{ g } HCap}{2.009_2 \times 10^{-3} \text{ mole } HCap} = 116.2 \text{ g/mole } HCap$$

The molar mass of caproic acid is 116.2 g/mole.

b. Caproic acid contains only C, H and O. If the elemental analysis of caproic acid indicates that it is 62.04% C, 10.41% H and 27.55% O by weight, what is the chemical formula of caproic acid?

Assume 100 g of caproic acid. *Note that not all work is shown; since it is identical to problems we've already seen. See me if you have questions on this problem.*

Element	%	Mass	Moles	Ratio of Moles Element to moles O
C	62.04	62.04	5.165 ₂₇	3.000
H	10.41	10.41	10.32 ₈₄	5.998
O	27.55	27.55	1.721 ₉₈	1.000

The empirical formula of caproic acid is C_3H_6O .

The empirical formula has a molar mass of 58.079 g/mole.

This molar mass is one half the known molecular mass, so the molecular formula is twice the empirical formula.

The molecular formula of caproic acid is $C_6H_{12}O_2$.

An alternate solution.

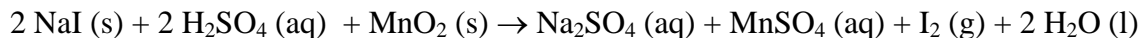
Assume 1.000 mole of caproic acid (i. e. 116.2 g/mole).

The mass of each element will be its percent multiplied by the molar mass.

The number of moles of each element in the compound is just the mass of each element divided by its atomic weight. This gives us the formula directly.

Element	%	Mass	Moles
C	62.04	72.09	6.002
H	10.41	12.09 ₆	12.00
O	27.55	32.01 ₃	2.001

7a. Gaseous I_2 may be prepared in the laboratory according to the following balanced chemical reaction.



What quantity of I_2 can be obtained if 20.0 g of NaI is mixed with 10.0 g MnO_2 and an excess of H_2SO_4 ? You are given that the molar mass of NaI is 149.89 g/mole, and that the molar mass of MnO_2 is 86.936 g/mole.

	2 NaI	2 H₂SO₄	MnO₂	→	I₂	Other products
Mass (g)	20.0	xs	10.0		16.9	
Molar mass (g/mole)	149.89	----	86.936		253.80	
moles	0.133		0.115		0.0665	

Determine moles of each reactant and limiting reagent.

$$20.0 \text{ g NaI} \left(\frac{1 \text{ mole NaI}}{149.89 \text{ g NaI}} \right) \left(\frac{1 \text{ mole I}_2}{2 \text{ mole NaI}} \right) \left(\frac{253.80 \text{ g I}_2}{1 \text{ mole I}_2} \right) = 16.9 \text{ g I}_2$$

$$10.0 \text{ g MnO}_2 \left(\frac{1 \text{ mole MnO}_2}{86.936 \text{ g MnO}_2} \right) \left(\frac{1 \text{ mole I}_2}{1 \text{ mole MnO}_2} \right) \left(\frac{253.80 \text{ g I}_2}{1 \text{ mole I}_2} \right) = 28.2 \text{ g I}_2$$

In this reaction NaI is the limiting reagent and 16.9 g I_2 are formed.

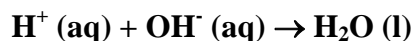
b. When the reaction in part a is performed, 10.0 g I_2 is recovered. What is the percent yield of I_2 ?

$$\% \text{ yield } I_2 = \left(\frac{\text{actual mass } I_2}{\text{theoretical mass } I_2} \right) \times 100 = \left(\frac{10.0 \text{ g } I_2}{16.9 \text{ g } I_2} \right) \times 100 = 59\%$$

The percent yield of I_2 in this reaction is 59%.

9. Suppose 124.0 mL of 0.129 M HCl is added to 101.0 mL of 0.151 M NaOH. Which ion is present in excess: H^+ or OH^- ? What is the final concentration of this ion (assume that the volumes are additive)? Show all work and justify your conclusions.

Start with balanced chemical equation. In this case, since both HCl and NaOH are strong electrolytes, we can use the net ionic equation:



	$H^+ (aq)$	+ $OH^- (aq)$	$\rightarrow H_2O (l)$
Volume (mL)	124.0	101.0	
Concentration (moles/L)	0.129	0.151	
Mass (grams)	-----	-----	
Molar mass (grams/mole)	-----	-----	
Moles	$1.60_0 \times 10^{-2}$	$1.52_5 \times 10^{-2}$	

Find moles of H^+ and moles of OH^- from volume of each reactant and their concentration.

$$124.0 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{0.129 \text{ mole } H^+}{1 \text{ L}} \right) = 1.60_0 \times 10^{-2} \text{ mole } H^+$$

$$101.0 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{0.151 \text{ mole } OH^-}{1 \text{ L}} \right) = 1.52_5 \times 10^{-2} \text{ mole } OH^-$$

Since the stoichiometry is 1:1 (one mole of H^+ reacts with one mole of OH^-), we can tell by inspection that there is more H^+ than OH^- . So, OH^- is the limiting reagent.

When the reaction is over we have no OH^- left and some H^+ left. The amount of H^+ left is (remember the amount of H^+ reacted equals amount of OH^- we had).

$$\text{final amount of } H^+ = \text{initial amount of } H^+ - \text{amount of } H^+ \text{ reacted}$$

$$\text{final amount of } H^+ = 1.60_0 \times 10^{-2} \text{ mole } H^+ - 1.52_5 \times 10^{-2} \text{ mole } H^+$$

$$\text{final amount of } H^+ = 7.5 \times 10^{-4} \text{ mole}$$

The concentration of H^+ is the amount of H^+ divided by the total volume.

$$[H^+] = \frac{7.5 \times 10^{-4} \text{ mole}}{124.0 \text{ mL} + 101.0 \text{ mL}} = \frac{7.5 \times 10^{-4} \text{ mole}}{225.0 \times 10^{-3} \text{ L}} = 3.3 \times 10^{-3} \text{ M}$$

The $[H^+] = 3.3 \times 10^{-3} \text{ M}$.

You could also take into account the amount of water formed by the reaction and add it to the total volume, but it makes a negligible contribution to the total volume.

10a. If 2.56 g of CaCO₃ are allowed to react with 250.0 mL of 0.125 M HCl, how much CaCl₂ will be formed? The balanced chemical reaction is



The molar masses are: CaCO₃, 100.09 g/mole; HCl, 36.46 g/mole; CaCl₂, 110.99 g/mole; CO₂, 44.01 g/mole and H₂O, 18.02 g/mole.

	CaCO ₃ (s) +	2 HCl (aq) →	CaCl ₂ (aq)	+ CO ₂ (g)	+ H ₂ O (l)
Mass (g)	2.56	-----		-----	-----
Molar mass (g/mole)	100.09	36.46	110.99	-----	-----
Concentration (M)	-----	0.125		-----	-----
Volume (mL)	-----	250.0		-----	-----
Moles				-----	-----

Determine limiting reagent.

$$2.56 \text{ g CaCO}_3 \left(\frac{1 \text{ mole CaCO}_3}{100.09 \text{ g CaCO}_3} \right) \left(\frac{1 \text{ mole CaCl}_2}{1 \text{ mole CaCO}_3} \right) \left(\frac{110.99 \text{ g CaCl}_2}{1 \text{ mole CaCl}_2} \right) = 2.84 \text{ g CaCl}_2$$

$$250.0 \times 10^{-3} \text{ L HCl} \left(\frac{0.125 \text{ mole HCl}}{1 \text{ L HCl}} \right) \left(\frac{1 \text{ mole CaCl}_2}{2 \text{ mole HCl}} \right) \left(\frac{110.99 \text{ g CaCl}_2}{1 \text{ mole CaCl}_2} \right) = 1.73 \text{ g CaCl}_2$$

HCl is the limiting reagent and so this reaction will form 1.73 g CaCl₂.

b. Assuming that the total volume at the end of the reaction is 250.0 mL, what will be the [Ca²⁺] when the reaction is over?

Moles of Ca²⁺ are calculated as follows (could also calculate from grams of CaCl₂)

$$0.03125 \text{ mole HCl} \left(\frac{1 \text{ mole CaCl}_2}{2 \text{ mole HCl}} \right) \left(\frac{1 \text{ mole Ca}^{2+}}{1 \text{ mole CaCl}_2} \right) = 0.015625 \text{ mole Ca}^{2+}$$

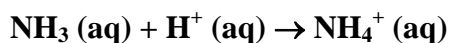
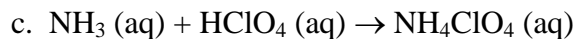
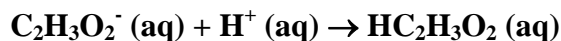
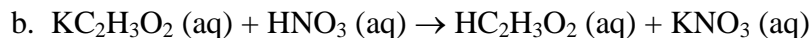
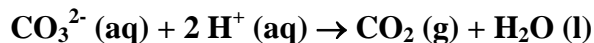
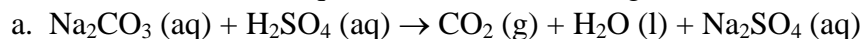
The concentration is then, by the definition of molarity,

$$\frac{0.015625 \text{ mole CaCl}_2}{250.0 \times 10^{-3} \text{ L}} = 0.0625 \text{ M}$$

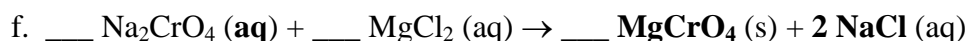
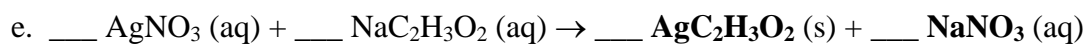
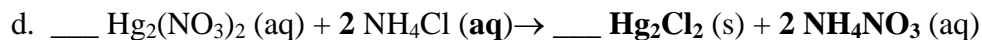
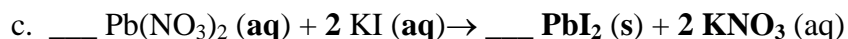
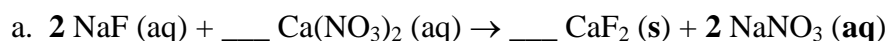
The [Ca²⁺] is 0.0625 M.

Note that the CaCO₃, although there is some left at the end of the reaction, does not make a significant contribution to the [Ca²⁺] because it is not soluble. We'll see later that this contribution is not zero!

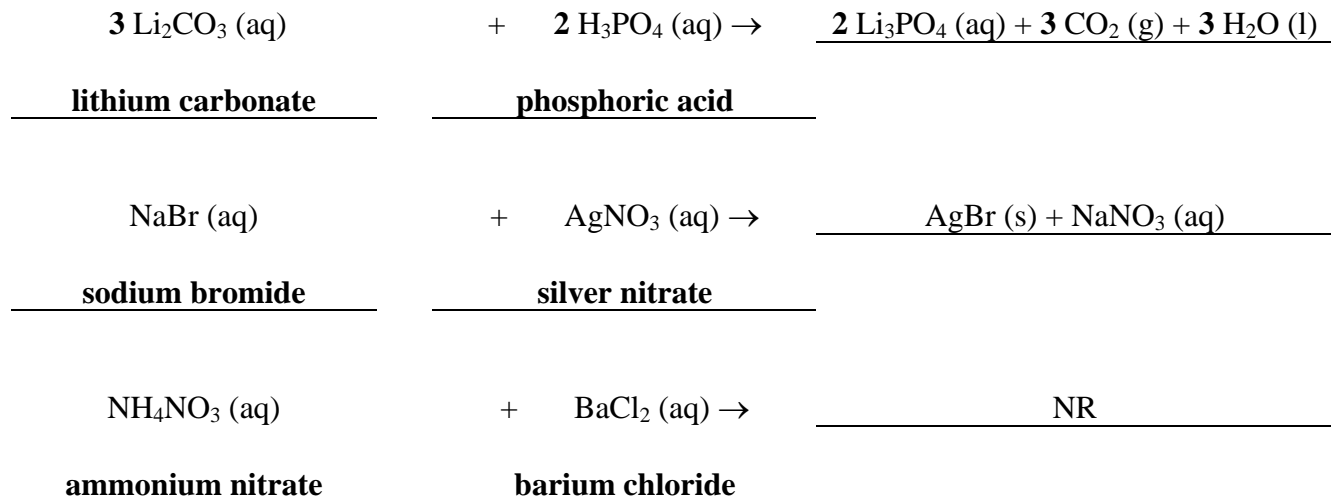
11. Give the net ionic equations for the following reactions.



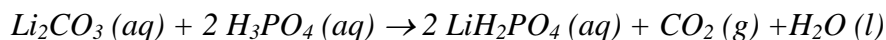
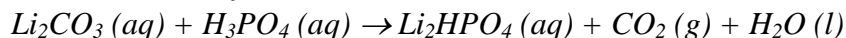
12. Fill in the blanks as appropriate.



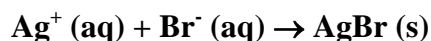
13a. Predict the products of the following reactions and balance the resulting chemical equations. If no reaction occurs, write "NR" on the product side. Be sure to indicate the state of each product. Write the name of each reactant in the space provided.



Alternate answers to the first chemical reaction:



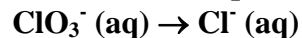
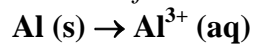
b. Write the net ionic equation for any two of the reactions in part a that give a reaction (e. g., if you entered NR to one of the reactions above, do not write a net ionic equation for it here).



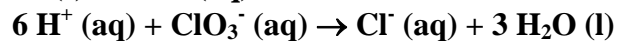
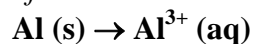
14. Balance the following redox reactions in either acidic or basic solution, as indicated. Identify the oxidant, the reductant what species is oxidized and which is reduced.

a. $\text{Al (s)} + \text{ClO}_3^- \text{ (aq)} \rightarrow \text{Cl}^- \text{ (aq)} + \text{Al}^{3+} \text{ (aq)}$ in acidic solution

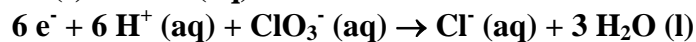
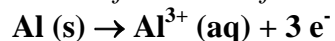
Separate into half-reactions.



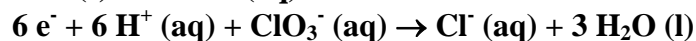
Balance for mass.



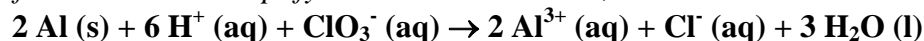
Balance the half-reactions for charge by adding electrons.



Balance electrons gained and lost by multiplying each half-reaction by a whole number.



Add half-reactions and simplify. Since we are in acid, we are done.

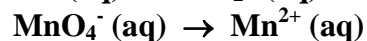
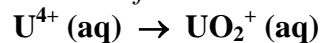


Al loses three electrons, so it is oxidized and it is the reductant.

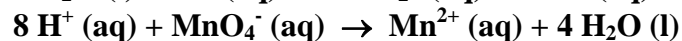
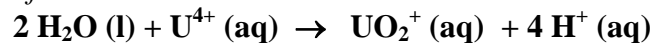
ClO₃⁻ gains six electrons, so it is reduced and it is the oxidant.

b. $\text{U}^{4+}(\text{aq}) + \text{MnO}_4^{-}(\text{aq}) \rightarrow \text{Mn}^{2+}(\text{aq}) + \text{UO}_2^{+}(\text{aq})$ in acidic solution

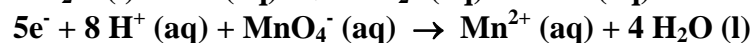
Separate into half-reactions.



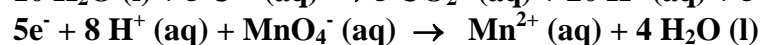
Balance for mass.



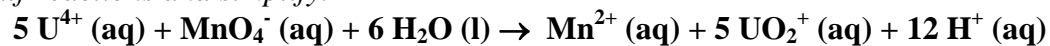
Balance the half-reactions for charge by adding electrons.



Balance electrons.



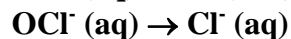
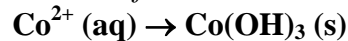
Add half-reactions and simplify.



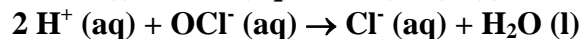
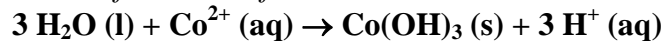
The oxidant is MnO_4^{-} (it is reduced) and the reductant is U^{4+} (it is oxidized).

c. $\text{Co}^{2+}(\text{aq}) + \text{OCl}^{-}(\text{aq}) \rightarrow \text{Co}(\text{OH})_3(\text{s}) + \text{Cl}^{-}(\text{aq})$ in basic solution

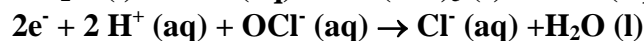
Separate into half-reactions.



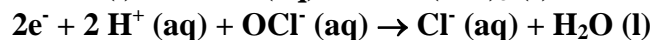
Balance the half-reactions for mass.



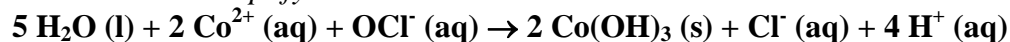
Balance the half-reactions for charge by adding electrons.



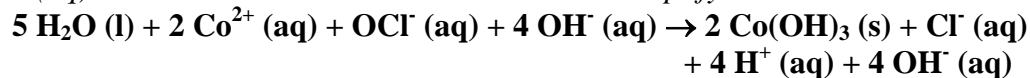
Balance the number of electrons.



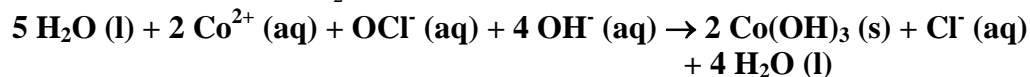
Add the reactions and simplify.



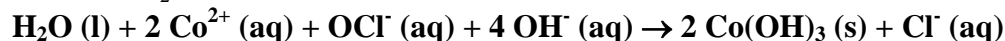
Add $\text{OH}^{-}(\text{aq})$ to both sides to “neutralize” the H^{+} and simplify.



Combine H^{+} and OH^{-} to make H_2O .



Cancel out extra H_2O .

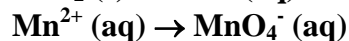
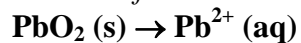


Co^{2+} loses an electron, and is oxidized, so it is the reductant

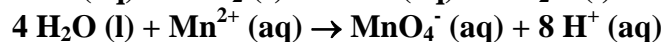
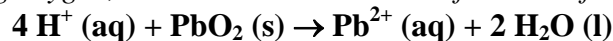
OCl^{-} gains two electrons, and is reduced, so it is the oxidant

d. $\text{PbO}_2 (\text{s}) + \text{Mn}^{2+} (\text{aq}) \rightarrow \text{Pb}^{2+} (\text{aq}) + \text{MnO}_4^- (\text{aq})$ in acidic solution

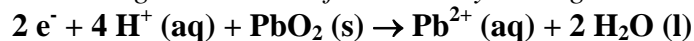
Separate into half reactions.



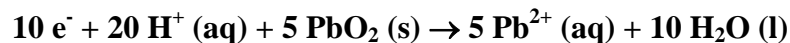
The half reactions are balanced for all elements except oxygen. Add H_2O to the side lacking oxygen, and H^+ to the other side of each half reaction to balance the oxygens.



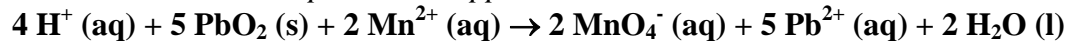
Balance the charges in each half reaction by adding electrons.



Balance the electrons.



Add the reactions and cancel species that appear on both sides.

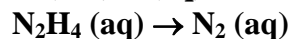
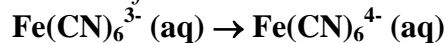


The oxidizing agent (oxidant) is PbO_2 ; it gets reduced.

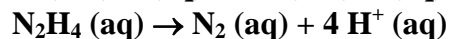
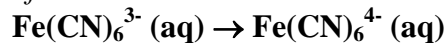
The reducing agent (reductant) is Mn^{2+} ; it gets oxidized.

e. $\text{Fe}(\text{CN})_6^{3-}(\text{aq}) + \text{N}_2\text{H}_4(\text{aq}) \rightarrow \text{Fe}(\text{CN})_6^{4-}(\text{aq}) + \text{N}_2(\text{g})$ in basic solution

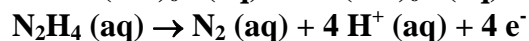
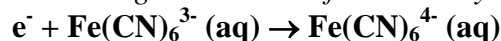
Separate into half reactions.



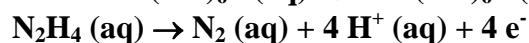
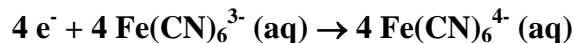
Balance for mass.



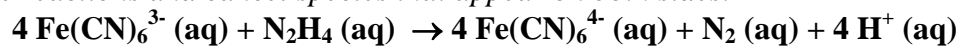
Balance the charges in each half reaction by adding electrons.



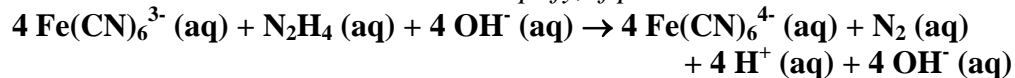
Balance the electrons.



Add the reactions and cancel species that appear on both sides.

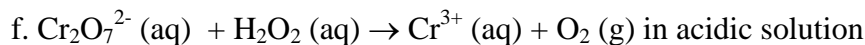


Add OH to both sides to "neutralize" H⁺. Simplify, if possible.

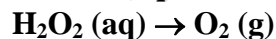
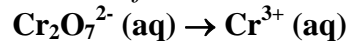


$\text{Fe}(\text{CN})_6^{3-}$ gains an electron so it is reduced, and it is the oxidant.

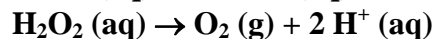
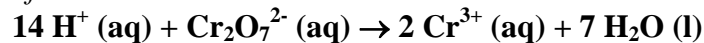
N_2H_4 loses four electrons so it is oxidized, and it is the reductant.



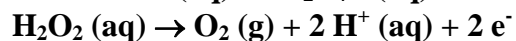
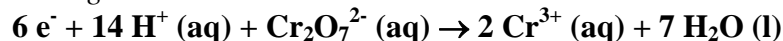
Separate into half-reactions.



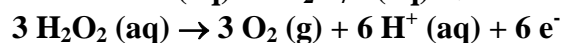
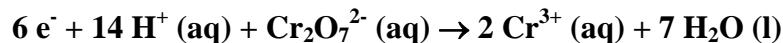
Balance for mass.



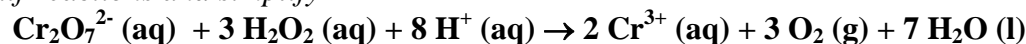
Balance charge.



Balance electrons.

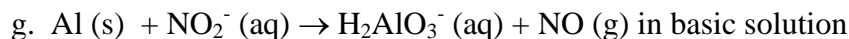


Add half-reactions and simplify

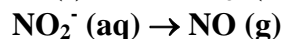


$\text{Cr}_2\text{O}_7^{2-}$ is the oxidant (it gains electrons, and loses oxygen, so it is reduced).

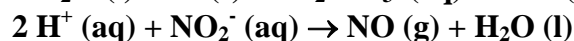
H_2O_2 is the reductant (it loses electrons, and loses hydrogen, so it is oxidized).



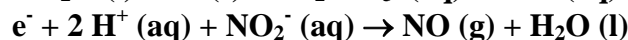
Separate into half-reactions.



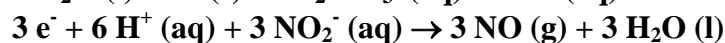
Balance for mass.



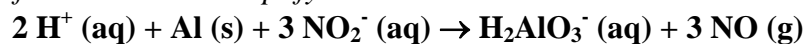
Balance charge.



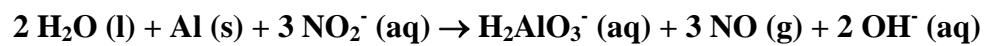
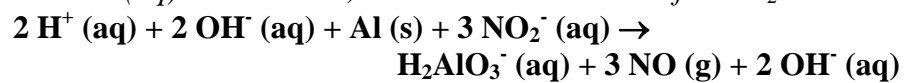
Balance electrons.



Add half-reactions and simplify



Now add 2 OH⁻ (aq) to both sides, combine H⁺ and OH⁻ to form H₂O and simplify.



**Al is oxidized (it loses electrons, and gains oxygens, and is the reducing agent).
NO₂⁻ is reduced (it gains electrons, and loses oxygen, and it is the oxidant)**