Chapter 3
Chemical Reactions

PRACTICING SKILLS

Balancing Equations

Balancing equations can be a matter of “running in circles” if a reasonable methodology is not employed. While there isn't one “right place” to begin, generally you will suffer fewer complications if you begin the balancing process using a substance that contains the **greatest number** of elements **or** the **largest subscript** values. Noting that you must have at least that many atoms of each element involved, coefficients can be used to increase the "atomic inventory". In the next few questions, you will see one **emboldened** substance in each equation. This emboldened substance is the one that I judge to be a “good” starting place. One last hint--modify the coefficients of uncombined elements, i.e. those not in compounds, <u>after</u> you modify the coefficients for compounds containing those elements -- <u>not before</u>!

1. Balanced equation for combustion of liquid pentane:
   $\mathbf{C_5H_{12}}$ (l) + 8 O_2 (g) → 6 H_2O (l) + 5 CO_2 (g)
   1. A minimum of 5 C and 12 H (a C_5H_{12} molecule) suggests coefficients of 5 for CO_2 and 6 for H_2O.
   2. Coefficients of 6 for H_2O and 5 for CO_2 will indicate a **total** of 16 O atoms or 8 molecules of the diatomic element, O_2.

3. (a) 4 Cr (s) + 3 O_2 (g) → 2 $\mathbf{Cr_2O_3}$ (s)
   1. Note the need for <u>at least</u> 2 Cr and 3 O atoms.
   2. Oxygen is diatomic -- we' ll need an <u>even</u> number of oxygen atoms, so try : 2 Cr_2O_3.
   3. 3 O_2 would give 6 O atoms on both sides of the equation.
   4. 4 Cr would give 4 Cr atoms on both sides of the equation.

   (b) Cu_2S (s) + O_2 (g) → □□Cu(s) + $\mathbf{SO_2}$ (g)
   1. A minimum of 2 O in SO_2 is required, and is provided with one molecule of elemental oxygen.
   2. 2 Cu atoms (on the right) indicates 2 Cu (on the left).

   (c) $\mathbf{C_6H_5CH_3}$ (l) + 9 O_2 (g) → 4 H_2O (l) + 7 CO_2 (g)
   1. A minimum of 7 C and 8 H is required.
   2. 7 CO_2 furnishes 7 C and 4 H_2O furnishes 8 H atoms.
   3. 4 H_2O and 7 CO_2 furnish a total of 18 O atoms, making the coefficient of O_2 = 9.

5. Balance and name the reactants and products:

(a) $\mathbf{Fe_2O_3}$(s) + 3 Mg(s) → 3 MgO(s) + 2 Fe(s)

1. Note the need for <u>at least</u> 2 Fe and 3 O atoms.
2. 2 Fe atoms would provide the proper iron atom inventory.
3. 3 MgO would give 3 O atoms on both sides of the equation.
4. 3 Mg would give 3 Mg atoms on both sides of the equation.

Reactants: iron(III) oxide and magnesium
Products: magnesium oxide and iron

(b) $\mathbf{AlCl_3}$(s) + 3 NaOH(aq) → $Al(OH)_3$(s) + 3 NaCl(aq)

1. Note the need for <u>at least</u> 1 Al and 3 Cl atoms.
2. 3 NaCl molecules would provide the proper Cl atom inventory.
3. 3 NaCl would require 3 Na atoms on the left side—a coefficient of 3 for NaOH is needed.
4. 3 OH groups (from $Al(OH)_3$) would give 3 OH groups needed on both sides of the equation—so a coefficient of 3 for NaOH is needed to provide that balance.

Reactants: aluminum chloride and sodium hydroxide
Products: aluminum hydroxide and sodium chloride.

(c) 2 $NaNO_3$(s) + H_2SO_4(l) → $\mathbf{Na_2SO_4}$(s) + 2 HNO_3(l)

1. Note the need for <u>at least</u> 2 Na and 1 S and 4 O atoms.
2. 2 $NaNO_3$ will provide the proper Na atom inventory.
3. The coefficient of 2 in front of $NaNO_3$ requires a coefficient of 2 for HNO_3—providing a balance for N atoms.
4. The implied coefficient of 1 for Na_2SO_4 suggests a similar coefficient for H_2SO_4—to balance the S atom inventory.
5. O atom inventory is done "automatically" when we balanced N and S inventories.

Reactants: sodium nitrate and sulfuric acid
Products: sodium sulfate and nitric acid
[....although nitric acid typically exists as an aqueous solution.]

(d) $\mathbf{NiCO_3}$(s) + 2 HNO_3(aq) → $Ni(NO_3)_2$(aq) + CO_2(g) + H_2O(l)

1. Note the need for <u>at least</u> 1 Ni atom on both sides. This inventory will mandate 2 NO_3 groups on the right—and also on the left. Since these come from HNO_3 molecules,we'll need 2 HNO_3 on the left.
2. The 2 H from the acid and the CO_3 from nickel carbonate, provide 2H, 1 C and 3 O atoms. 1 H_2O takes care of the 2H, and **one** of the O atoms, 1 CO_2 consumes the 1 C and the remaining 2 O atoms.

Reactants: nickel(II) carbonate and nitric acid
Products: nickel(II) nitrate, carbon dioxide, and water

Chemical Equilibrium

7. The greater electrical conductivity of the HCl solution at equilibrium indicates a greater concentration of ions (H_3O^+ and Cl^-), indicating that the HCl solution is more product-favored at equilibrium than the HCO_2H solution.

Ions and Molecules in Aqueous Solution

9. What is an electrolyte? What are experimental means for discriminating between weak and strong electrolytes?
An electrolyte is a substance whose aqueous solution conducts an electric current.
As to experimental means for discriminating between weak and strong electrolytes, refer to the apparatus in the Active Figure5.2. NaCl is a strong electrolyte and would cause the bulb to glow brightly—reflecting a large number of ions in solution while aqueous ammonia or vinegar (an aqueous solution of acetic acid) would cause the bulb to glow only dimly—indicating a smaller number of ions in solution.

11. Predict water solubility:
(a) $CuCl_2$ is expected to be soluble, while CuO and $FeCO_3$ are not. Chlorides are generally water soluble, while oxides and carbonates are not.
(b) $AgNO_3$ is soluble. AgI and Ag_3PO_4 are not soluble. Nitrate salts are soluble. Phosphate salts are generally insoluble. While halides are generally soluble, those of Ag^+ are not.
(c) K_2CO_3, KI and $KMnO_4$ are soluble. In general, salts of the alkali metals are soluble.

13. Ions produced when the compounds dissolve in water.

Compound	Cation	Anion
(a) KOH	K^+	OH^-
(b) K_2SO_4	2 K^+	SO_4^{2-}
(c) $LiNO_3$	Li^+	NO_3^-
(d) $(NH_4)_2SO_4$	2 NH_4^+	SO_4^{2-}

15.

Compound	Water Soluble	Cation	Anion
(a) Na_2CO_3	yes	2 Na^+	CO_3^{2-}
(b) $CuSO_4$	yes	Cu^{2+}	SO_4^{2-}
(c) NiS	no		
(d) $BaBr_2$	yes	Ba^{2+}	2 Br^-

Precipitation Reactions and Net Ionic Equations

17. $CdCl_2(aq) + 2\,NaOH(aq) \rightarrow Cd(OH)_2(s) + 2\,NaCl(aq)$

    Net ionic equation: $Cd^{2+}(aq) + 2\,OH^-(aq) \rightarrow Cd(OH)_2(s)$

19. Balanced equations for precipitation reactions:

    (a) $NiCl_2(aq) + (NH_4)_2S(aq) \rightarrow NiS(s) + 2\,NH_4Cl(aq)$

    Net ionic equation: $Ni^{2+}(aq) + S^{2-}(aq) \rightarrow NiS(s)$

    (b) $3\,Mn(NO_3)_2(aq) + 2\,Na_3PO_4(aq) \rightarrow Mn_3(PO_4)_2(s) + 6\,NaNO_3(aq)$

    Net ionic equation: $3\,Mn^{2+}(aq) + 2\,PO_4{}^{3-}(aq) \rightarrow Mn_3(PO_4)_2(s)$

Acids and Bases and Their Reactions

21. $HNO_3(aq) + H_2O$(Ú) $\rightarrow H_3O^+(aq) + NO_3{}^-(aq)$

    alternatively: $HNO_3(aq) \rightarrow H^+(aq) + NO_3{}^-(aq)$

23. $H_2C_2O_4(aq) \rightarrow H^+(aq) + HC_2O_4{}^-(aq)$

    $HC_2O_4{}^-(aq) \rightarrow H^+(aq) + C_2O_4{}^{2-}(aq)$

25. $MgO(s) + H_2O(\ell) \rightarrow Mg(OH)_2(s)$ (metal oxide reacts with water to form a base)

27. Complete and Balance

    (a) $2\,CH_3CO_2H(aq) + Mg(OH)_2(s) \rightarrow Mg(CH_3CO_2)_2(aq) + 2\,H_2O(\ell)$

acetic acid	magnesium hydroxide(base)	magnesium acetate	water

    (b) $HClO_4(aq) + NH_3(aq) \rightarrow NH_4ClO_4(aq)$

perchloric acid	ammonia (base)	ammonium perchlorate

29. Write and balance the equation for barium hydroxide reacting with nitric acid:

    $Ba(OH)_2(s) + 2\,HNO_3(aq) \rightarrow Ba(NO_3)_2(aq) + 2\,H_2O$(Ú)

barium hydroxide	nitric acid	barium nitrate	water

31. Two strong Brönsted acids and one strong Brönsted base:

    Many examples exist: Strong acids: HCl, HBr, HI, HNO_3

    Strong bases: LiOH, NaOH, KOH

Writing Net Ionic Equations

33. (a) $(NH_4)_2CO_3(aq) + Cu(NO_3)_2(aq) \rightarrow CuCO_3(s) + 2\ NH_4NO_3(aq)$

(net) $CO_3^{2-}(aq) + Cu^{2+}(aq) \rightarrow CuCO_3(s)$

(b) $Pb(OH)_2(s) + 2\ HCl(aq) \rightarrow PbCl_2(s) + 2\ H_2O(l)$

(net) $Pb(OH)_2(s) + 2\ H_3O^+(aq) + 2\ Cl^-(aq) \rightarrow PbCl_2(s) + 4\ H_2O(l)$

(c) $BaCO_3(s) + 2\ HCl(aq) \rightarrow BaCl_2(aq) + H_2O(l) + CO_2(g)$

(net) $BaCO_3(s) + 2\ H^+(aq) \rightarrow Ba^{2+}(aq) + H_2O(l) + CO_2(g)$

alternatively: $BaCO_3(s) + 2\ H_3O^+(aq) \rightarrow Ba^{2+}(aq) + 3\ H_2O(l) + CO_2(g)$

(d) $2\ CH_3CO_2H(aq) + Ni(OH)_2(s) \rightarrow Ni(CH_3CO_2)_2(aq) + H_2O(l)$

net: $2\ CH_3CO_2H(aq) + Ni(OH)_2(s) \rightarrow Ni^{2+}(aq) + 2\ CH_3CO_2^-(aq) + 2\ H_2O(l)$

35. (a) $AgNO_3(aq) + KI(aq) \rightarrow AgI(s) + KNO_3(aq)$

(net) $Ag^+(aq) + I^-(aq) \rightarrow AgI(s)$

(b) $Ba(OH)_2(aq) + 2\ HNO_3(aq) \rightarrow 2\ H_2O(l) + Ba(NO_3)_2(aq)$

(net) $OH^-(aq) + H_3O^+(aq) \rightarrow 2\ H_2O(l)$

(c) $2\ Na_3PO_4(aq) + 3\ Ni(NO_3)_2(aq) \rightarrow Ni_3(PO_4)_2(s) + 6\ NaNO_3(aq)$

(net) $2\ PO_4^{3-}(aq) + 3\ Ni^{2+}(aq) \rightarrow Ni_3(PO_4)_2(s)$

Gas-Forming Reactions

37. Write and balance the equation for iron(II) carbonate reacting with nitric acid:

$FeCO_3(s) + 2\ HNO_3(aq) \rightarrow Fe(NO_3)_2(aq) + H_2O(l) + CO_2(g)$

iron(II) carbonate	nitric acid	iron(II) nitrate	water	carbon dioxide

39. Overall, balanced equation for reaction of $(NH_4)_2S$ with HBr:

$(NH_4)_2S$	+ 2 HBr	$\rightarrow$	H_2S	+ 2 NH_4Br
Ammonium sulfide	Hydrogen bromide		Hydrogen sulfide	Ammonium bromide

Oxidation Numbers

41. For questions on oxidation number, read the symbol (x) as "the oxidation number of x."

(a) BrO_3^- $\quad$ (Br) + 3(O) = -1

Since oxygen almost always has an oxidation number of -2, we can substitute this value and solve for the oxidation number of Br.

(Br) + 3(-2) = -1

(Br) = +5

(b) $C_2O_4^{2-}$

$2\,(C) + 4\,(O) = -2$

$2\,(C) + 4\,(-2) = -2$

$2\,(C) + -8 = -2$

$2\,(C) = +6$

$(C) = +3$

(c) F^- The oxidation number for any monatomic ion is the charge on the ion. So (F) = -1

(d) CaH_2

$(Ca) + 2\,(H) = 0$

$(Ca) + 2\,(-1) = 0$

$(Ca) = +2$

(e) H_4SiO_4

$4(H) + (Si) + 4(O) = 0$

$4(+1) + (Si) + 4(-2) = 0$

$(Si) = +4$

(f) HSO_4^-

$(H) + (S) + 4(O) = -1$

$(+1) + (S) + 4(-2) = -1$

$(S) = +6$

Oxidation-Reduction Reactions

43. (a) Oxidation-Reduction: Zn(s) has an oxidation number of 0, while Zn^{2+}(aq) has an oxidation number of +2—hence Zn is being oxidized.
N in NO_3^- has an oxidation number of +5, while N in NO_2 has an oxidation number of +4—hence N is being reduced.

(b) Acid-Base reaction: There is no change in oxidation number for any of the elements in this reaction—hence it is NOT an oxidation-reduction reaction.
H_2SO_4 is an acid, and $Zn(OH)_2$ acts as a base.

(c) Oxidation-Reduction: Ca(s) has an oxidation number of 0, while Ca^{2+}(aq) has an oxidation number of +2—hence Ca is being oxidized.
H in H_2O has an oxidation number of +1, while H in H_2 has an oxidation number of 0—hence H is being reduced.

45. Determine which reactant is oxidized and which is reduced:

(a) $C_2H_4(g) + 3\ O_2(g) \rightarrow 2CO_2(g) + 2\ H_2O(g)$

specie	ox. number before	after	has experienced	functions as the
C	-2	+4	oxidation	(C_2H_4) reducing agent
H	+1	+1	no change	
O	0	-2	reduction	(O_2) oxidizing agent

(b) $Si(s) + 2\ Cl_2(g) \rightarrow SiCl_4(l)$

specie	ox. number before	after	has experienced	functions as the
Si	0	+4	oxidation	(Si) reducing agent
Cl	0	-1	reduction	(Cl_2) oxidizing agent

Types of Reactions in Aqueous Solution

47. Precipitation (PR), Acid-Base (AB), or Gas-Forming (GF)

(a) $Ba(OH)_2(s) + 2\ HCl(aq) \rightarrow BaCl_2(aq) + 2\ H_2O(l)$ AB

(b) $2\ HNO_3(aq) + CoCO_3(s) \rightarrow Co(NO_3)_2(aq) + H_2O(l) + CO_2(g)$ GF

(c) $2\ Na_3PO_4(aq) + 3\ Cu(NO_3)_2(aq) \rightarrow Cu_3(PO_4)_2(s) + 6\ NaNO_3(aq)$ PR

49. Precipitation (PR), Acid-Base (AB), or Gas-Forming (GF)

(a) $MnCl_2(aq) + Na_2S(aq) \rightarrow MnS(s) + 2\ NaCl(aq)$ PR

(net) $Mn^{2+}(aq) + S^{2-}(aq) \rightarrow MnS(s)$

(b) $K_2CO_3(aq) + ZnCl_2(aq) \rightarrow ZnCO_3(s) + 2\ KCl(aq)$ PR

(net) $CO_3^{2-}(aq) + Zn^{2+}(aq) \rightarrow ZnCO_3(s)$

51. Balance the following and classify them as PR, AB, GF, or OR:

(a) $CuCl_2\ (aq) + H_2S(aq) \rightarrow CuS(s) + 2\ HCl(aq)$ PR

(b) $H_3PO_4\ (aq) + 3\ KOH(aq) \rightarrow 3\ H_2O(l) + K_3PO_4(aq)$ AB

(c) $Ca(s) + 2\ HBr(aq) \rightarrow H_2(g) + CaBr_2(aq)$ OR

(d) $MgCl_2(aq) + H_2O(l) \rightarrow 3\ Mg(OH)_2(s) + 2\ HCl(aq)$ PR

GENERAL QUESTIONS

53. Balance:

(a) Synthesis of urea:

$CO_2(g) + 2\ NH_3(g) \rightarrow \mathbf{CO(NH_2)_2}(s) + H_2O(l)$

1. Note the need for two NH_3 in each molecule of urea, so multiply NH_3 by 2.
2. 2 NH_3 provides the two H atoms for a molecule of H_2O.
3. Each CO_2 provides the O atom for a molecule of H_2O.

(b) synthesis of uranium(VI) fluoride

$UO_2(s) + 4\,HF(aq) \rightarrow \mathbf{UF_4}(s) + 2\,H_2O(l)$

$UF_4(s) + F_2(g) \rightarrow \mathbf{UF_6}(s)$

1. The 4 F atoms in UF_4 requires 4 F atoms from HF. (equation 1)
2. The H atoms in HF produce 2 molecules of H_2O. (equation 1)
3. The 1:1 stoichiometry of $UF_6 : UF_4$ provides a simple balance. (equation 2)

(c) synthesis of titanium metal from TiO_2:

$\mathbf{TiO_2}(s) + 2\,Cl_2(g) + 2\,C(s) \rightarrow TiCl_4(l) + 2\,CO(g)$

$\mathbf{TiCl_4}(l) + 2\,Mg(s) \rightarrow Ti(s) + 2\,MgCl_2(s)$

1. The O balance mandates 2 CO for each TiO_2. (equation 1)
2. A coefficient of 2 for C provides C balance. (equation 1)
3. The Ti balance (TiO_2:$TiCl_4$) requires 4 Cl atoms, hence 2 Cl_2 (equation 1)
4. The Cl balance requires 2 $MgCl_2$, hence 2 Mg. (equation 2)

55. Formula for the following compounds:

(a) soluble compound with Br^- ion: almost any bromide compound with the exception of Ag^+, Hg_2^{2+} and Pb^{2+}

(b) insoluble hydroxide: almost any hydroxide except salts of NH_4^+ and the alkali metal ions

(c) insoluble carbonate: almost any carbonate except salts of NH_4^+ and the alkali metal ions

(d) soluble nitrate-containing compound: all nitrate-containing compounds are soluble
The listing of soluble and insoluble compounds in your text will provide general guidelines for predicting the solubility of compounds.

(e) a weak Bronsted acid: the carboxylic acids are weak acids: CH_3CO_2H (acetic)

57. For the following copper salts:

Water soluble: $Cu(NO_3)_2$, $CuCl_2$ — nitrates and chlorides are soluble

Water insoluble: $CuCO_3$, $Cu_3(PO_4)_2$ — carbonates and phosphates are insoluble

59. **Spectator ions** in the following equation and the net ionic equation:

$2\,H_3O^+(aq) + 2\,\mathbf{NO_3^-(aq)} + Mg(OH)_2(s) \rightarrow 4\,H_2O(l) + Mg^{2+}(aq) + 2\,\mathbf{NO_3^-(aq)}$

The emboldened nitrate ions are the spectator ions. The net ionic equation would be the first equation shown above without the spectator ions:

$2\,H_3O^+(aq) + Mg(OH)_2(s) \rightarrow 4\,H_2O(l) + Mg^{2+}(aq)$ [An acid-base exchange]

61. For the reaction of chlorine with NaBr: $Cl_2(g) + 2\ NaBr(aq) \rightarrow 2\ NaCl(aq) + Br_2(l)$

(a) Oxidized: **bromine's** oxidation number is changed from -1 to 0

Reduced: **chlorine's** oxidation number is changed from 0 to –1

(b) Oxidizing agent: **Cl_2** removes the electrons from NaBr

Reducing agent: **NaBr** provides the electrons to the chlorine.

63. Reaction: $MgCO_3(s) + 2\ HCl(aq) \rightarrow CO_2(g) + MgCl_2(aq) + H_2O(l)$

(a) The net ionic equation: $MgCO_3(s) + 2\ H_3O^+(aq) \rightarrow CO_2(g) + Mg^{2+}(aq) + 3\ H_2O(l)$

The spectator ion is the chloride ion (Cl^-).

(b) The production of $CO_2(g)$ characterizes this as a gas-forming reaction.

65. Species present in aqueous solutions of:

	compound	types of species	species present
(a)	NH_3	molecules (weak base)	NH_3, NH_4^+, OH^-
(b)	CH_3CO_2H	molecules (weak acid)	CH_3CO_2H, $CH_3CO_2^-$,H^+
(c)	NaOH	ions (strong base)	Na^+ and OH^-
(d)	HBr	ions (strong acid)	H_3O^+ and Br^-

In every case, H_2O will be present (but omitted in this list)

67. Balance and classify each as PR, AB, GF

(a)	$K_2CO_3(aq) + 2\ HClO_4(aq) \rightarrow 2\ KClO_4(aq) + CO_2(g) + 2\ H_2O(l)$	GF
(b)	$FeCl_2(aq) + (NH_4)_2S(aq) \rightarrow FeS(s) + 2\ NH_4Cl(aq)$	PR
(c)	$Fe(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow FeCO_3(s) + 2\ NaNO_3(aq)$	PR
(d)	$3\ NaOH(aq) + FeCl_3(aq) \rightarrow 3\ NaCl(aq) + Fe(OH)_3(s)$	PR

IN THE LABORATORY

69. For the reaction:

$2\ NaI(s) + 2\ H_2SO_4(aq) + MnO_2(s) \rightarrow Na_2SO_4(aq) + MnSO_4(aq) + I_2\ (g) + 2\ H_2O(l)$

(a) Oxidation number of each atom in the equation: (ox.numbers shown in order)

Reactants: NaI (+1,-1) H_2SO_4 (+1, +6, -2) MnO_2 (+4,-2)

Products: Na_2SO_4 (+1,+6,-2) $MnSO_4$ (+2,+6,-2) I_2(0) H_2O(+1,-2)

(b) Oxidizing agent: MnO_2 Oxidized: I in NaI

Reducing agent: NaI Reduced: Mn (in MnO_2)

(c) The formation of gaseous iodine "drives" the process—product-favored

(d) Names of reactants and products:

NaI	H_2SO_4	MnO_2	Na_2SO_4	$MnSO_4$	I_2	H_2O
sodium iodide	sulfuric acid	manganese(IV) oxide	sodium sulfate	manganese(II) sulfate	iodine	water

71. Another way to prepare $MgCl_2$: Given the reactivity of both elemental magnesium and chlorine, one can bring the two elements into contact (carefully!)

$$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$$

73. In the reaction:

$C_6H_{12}O_6$ (aq) + 2 Ag^+ (aq) + 2 OH^- (aq) → $C_6H_{12}O_7$ (aq) + 2 Ag(s) + H_2O(l)

Oxidized: $C_6H_{12}O_6$ is oxidized to $C_6H_{12}O_7$ (simple observation—note that O is added)

Reduced: Ag^+(aq) is reduced to Ag(s) (oxidation number changes from +1 to 0)

Oxidizing agent: Ag^+ (aq) oxidizes the sugar

Reducing agent: $C_6H_{12}O_6$ reduces Ag^+

SUMMARY AND CONCEPTUAL QUESTIONS

75. A simple experiment to prove that lactic acid is a weak acid (ionizing to a small extent) is to test the conductivity of the solution. A conductivity apparatus (e.g. a light bulb) will indicate only a small current flow (a light bulb will glow only dimly).

To prove that the establishment of equilibrium is reversible, add strong acid (H_3O^+). The shift of equilibrium to the left should result in the molecular acid precipitating from solution.

77. Using the reagents: $BaCl_2$, $BaCO_3$, $Ba(OH)_2$, H_2SO_4, Na_2SO_4,

Prepare barium sulfate by:

a precipitation reaction

The reaction of $BaCl_2$ with Na_2SO_4 will perform this task:

$BaCl_2$ (aq) + Na_2SO_4(aq) → $BaSO_4$(s) + 2 NaCl (aq)

a gas-forming reaction

$BaCO_3$(s) + H_2SO_4(l) → $BaSO_4$(s) + H_2O(l) + CO_2(g)

One might think about using $Ba(OH)_2$ as one reactant for part (a), but the substance isn't very water-soluble. $BaCl_2$ is much more water-soluble.