

Chapter 3

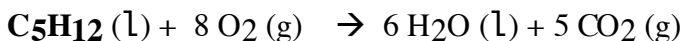
Chemical Reactions

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, after you modify the coefficients for compounds containing those elements -- not before!

1. Balanced equation for combustion of liquid pentane:



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) $4 Cr (s) + 3 O_2 (g) \rightarrow 2 \mathbf{Cr_2O_3} (s)$

1. Note the need for at least 2 Cr and 3 O atoms.
2. Oxygen is diatomic -- we'll need an even 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.

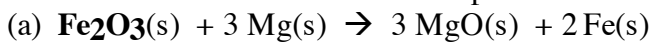
(b) $Cu_2S (s) + O_2 (g) \rightarrow \square\square Cu(s) + \mathbf{SO_2} (g)$

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) $\mathbf{C_6H_5CH_3} (l) + 9 O_2 (g) \rightarrow 4 H_2O (l) + 7 CO_2 (g)$

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 $O_2 = 9$.

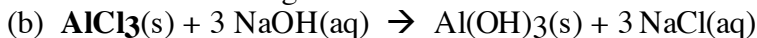
5. Balance and name the reactants and products:



1. Note the need for at least 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

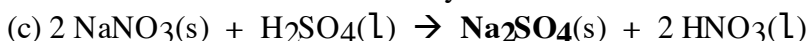
Products: magnesium oxide and iron



1. Note the need for at least 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 $\text{Al}(\text{OH})_3$) 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.

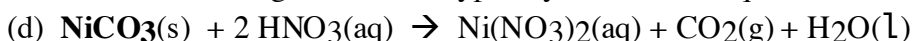


1. Note the need for at least 2 Na and 1 S and 4 O atoms.
2. 2 NaNO_3 will provide the proper Na atom inventory.
3. The coefficient of 2 in front of NaNO_3 requires a coefficient of 2 for HNO_3 —providing a balance for N atoms.
4. The implied coefficient of 1 for Na_2SO_4 suggests a similar coefficient 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.]



1. Note the need for at least 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 Figure 5.2. 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.

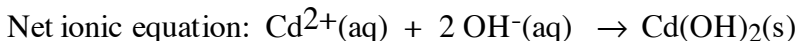
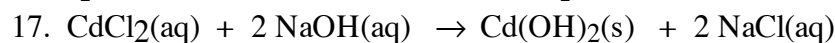
11. Predict water solubility:

- (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.

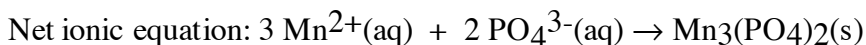
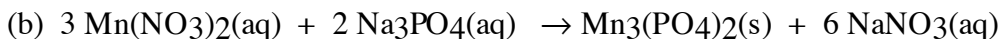
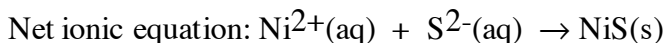
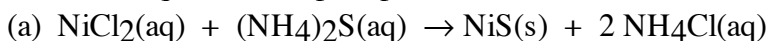
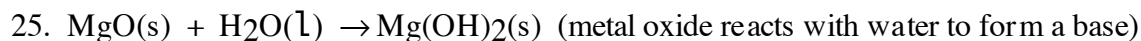
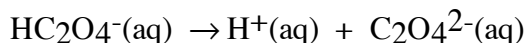
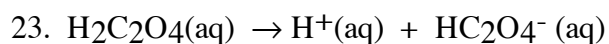
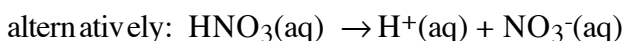
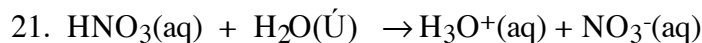
13. Ions produced when the compounds dissolve in water.

<u>Compound</u>	<u>Cation</u>	<u>Anion</u>
(a) KOH	K^+	OH^-
(b) K_2SO_4	2K^+	SO_4^{2-}
(c) LiNO_3	Li^+	NO_3^-
(d) $(\text{NH}_4)_2\text{SO}_4$	2NH_4^+	SO_4^{2-}

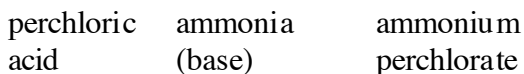
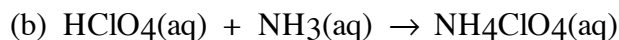
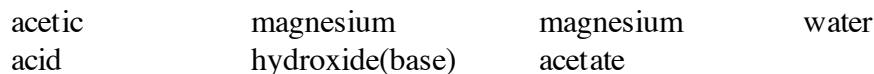
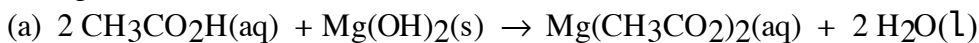
<u>Compound</u>	<u>Water Soluble</u>	<u>Cation</u>	<u>Anion</u>
(a) Na_2CO_3	yes	2Na^+	CO_3^{2-}
(b) CuSO_4	yes	Cu^{2+}	SO_4^{2-}
(c) NiS	no		
(d) BaBr_2	yes	Ba^{2+}	2Br^-

Precipitation Reactions and Net Ionic Equations

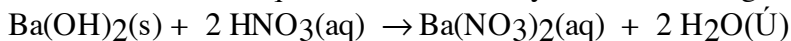
19. Balanced equations for precipitation reactions:

**Acids and Bases and Their Reactions**

27. Complete and Balance



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



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

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

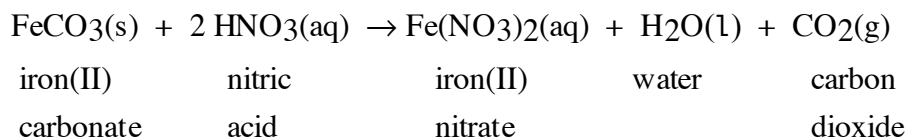
Strong bases: LiOH , NaOH , KOH

Writing Net Ionic Equations

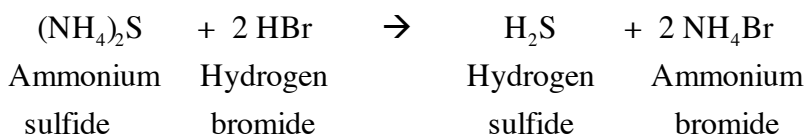
33. (a) $(\text{NH}_4)_2\text{CO}_3(\text{aq}) + \text{Cu}(\text{NO}_3)_2(\text{aq}) \rightarrow \text{CuCO}_3(\text{s}) + 2 \text{NH}_4\text{NO}_3(\text{aq})$
 (net) $\text{CO}_3^{2-}(\text{aq}) + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{CuCO}_3(\text{s})$
- (b) $\text{Pb}(\text{OH})_2(\text{s}) + 2 \text{HCl}(\text{aq}) \rightarrow \text{PbCl}_2(\text{s}) + 2 \text{H}_2\text{O}(\text{l})$
 (net) $\text{Pb}(\text{OH})_2(\text{s}) + 2 \text{H}_3\text{O}^+(\text{aq}) + 2 \text{Cl}^-(\text{aq}) \rightarrow \text{PbCl}_2(\text{s}) + 4 \text{H}_2\text{O}(\text{l})$
- (c) $\text{BaCO}_3(\text{s}) + 2 \text{HCl}(\text{aq}) \rightarrow \text{BaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
 (net) $\text{BaCO}_3(\text{s}) + 2 \text{H}^+(\text{aq}) \rightarrow \text{Ba}^{2+}(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
 alternatively: $\text{BaCO}_3(\text{s}) + 2 \text{H}_3\text{O}^+(\text{aq}) \rightarrow \text{Ba}^{2+}(\text{aq}) + 3 \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g})$
- (d) $2 \text{CH}_3\text{CO}_2\text{H}(\text{aq}) + \text{Ni}(\text{OH})_2(\text{s}) \rightarrow \text{Ni}(\text{CH}_3\text{CO}_2)_2(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 net: $2 \text{CH}_3\text{CO}_2\text{H}(\text{aq}) + \text{Ni}(\text{OH})_2(\text{s}) \rightarrow \text{Ni}^{2+}(\text{aq}) + 2 \text{CH}_3\text{CO}_2^-(\text{aq}) + 2 \text{H}_2\text{O}(\text{l})$
35. (a) $\text{AgNO}_3(\text{aq}) + \text{KI}(\text{aq}) \rightarrow \text{AgI}(\text{s}) + \text{KNO}_3(\text{aq})$
 (net) $\text{Ag}^+(\text{aq}) + \text{I}^-(\text{aq}) \rightarrow \text{AgI}(\text{s})$
- (b) $\text{Ba}(\text{OH})_2(\text{aq}) + 2 \text{HNO}_3(\text{aq}) \rightarrow 2 \text{H}_2\text{O}(\text{l}) + \text{Ba}(\text{NO}_3)_2(\text{aq})$
 (net) $\text{OH}^-(\text{aq}) + \text{H}_3\text{O}^+(\text{aq}) \rightarrow 2 \text{H}_2\text{O}(\text{l})$
- (c) $2 \text{Na}_3\text{PO}_4(\text{aq}) + 3 \text{Ni}(\text{NO}_3)_2(\text{aq}) \rightarrow \text{Ni}_3(\text{PO}_4)_2(\text{s}) + 6 \text{NaNO}_3(\text{aq})$
 (net) $2 \text{PO}_4^{3-}(\text{aq}) + 3 \text{Ni}^{2+}(\text{aq}) \rightarrow \text{Ni}_3(\text{PO}_4)_2(\text{s})$

Gas-Forming Reactions

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



39. Overall, balanced equation for reaction of $(\text{NH}_4)_2\text{S}$ with HBr:

**Oxidation Numbers**

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

$$\text{(a) } \text{BrO}_3^- \quad (\text{Br}) + 3(\text{O}) = -1$$

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

$$(\text{Br}) + 3(-2) = -1$$

$$(\text{Br}) = +5$$

$$\begin{aligned}
 \text{(b) } \text{C}_2\text{O}_4^{2-} & \quad 2(\text{C}) + 4(\text{O}) = -2 \\
 & \quad 2(\text{C}) + 4(-2) = -2 \\
 & \quad 2(\text{C}) + -8 = -2 \\
 & \quad 2(\text{C}) = +6 \\
 & \quad (\text{C}) = +3
 \end{aligned}$$

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

$$\begin{aligned}
 \text{(d) } \text{CaH}_2 & \quad (\text{Ca}) + 2(\text{H}) = 0 \\
 & \quad (\text{Ca}) + 2(-1) = 0 \\
 & \quad (\text{Ca}) = +2
 \end{aligned}$$

$$\begin{aligned}
 \text{(e) } \text{H}_4\text{SiO}_4 & \quad 4(\text{H}) + (\text{Si}) + 4(\text{O}) = 0 \\
 & \quad 4(+1) + (\text{Si}) + 4(-2) = 0 \\
 & \quad (\text{Si}) = +4
 \end{aligned}$$

$$\begin{aligned}
 \text{(f) } \text{HSO}_4^- & \quad (\text{H}) + (\text{S}) + 4(\text{O}) = -1 \\
 & \quad (+1) + (\text{S}) + 4(-2) = -1 \\
 & \quad (\text{S}) = +6
 \end{aligned}$$

Oxidation-Reduction Reactions

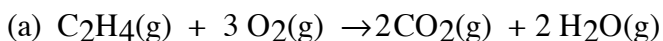
43. (a) Oxidation-Reduction: $\text{Zn}(\text{s})$ has an oxidation number of 0, while $\text{Zn}^{2+}(\text{aq})$ has an oxidation number of +2—hence Zn is being oxidized. N in NO_3^- has an oxidation number of +5, while N in NO_2 has an oxidation number of +4—hence N is being reduced.

(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.

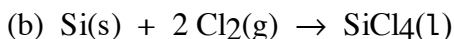
H_2SO_4 is an acid, and $\text{Zn}(\text{OH})_2$ acts as a base.

(c) Oxidation-Reduction: $\text{Ca}(\text{s})$ has an oxidation number of 0, while $\text{Ca}^{2+}(\text{aq})$ 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:



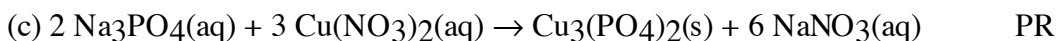
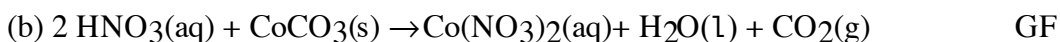
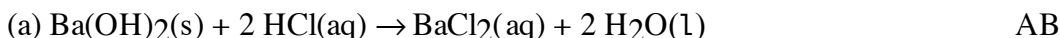
		ox. number			
specie	before	after	has experienced	functions as the	
C	-2	+4	oxidation	(C ₂ H ₄)	reducing agent
H	+1	+1	no change		
O	0	-2	reduction	(O ₂)	oxidizing agent



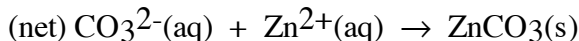
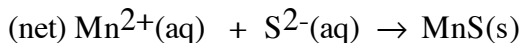
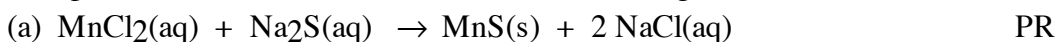
		ox. number			
specie	before	after	has experienced	functions as the	
Si	0	+4	oxidation	(Si)	reducing agent
Cl	0	-1	reduction	(Cl ₂)	oxidizing agent

Types of Reactions in Aqueous Solution

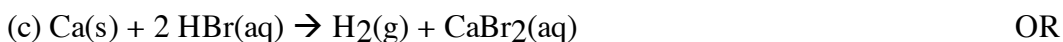
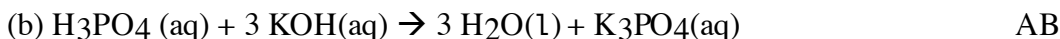
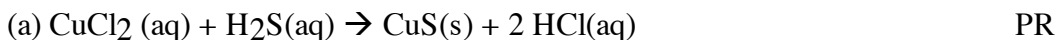
47. Precipitation (PR), Acid-Base (AB), or Gas-Forming (GF)



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



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



GENERAL QUESTIONS

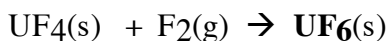
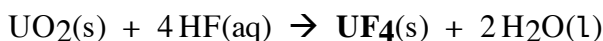
53. Balance:

(a) Synthesis of urea:



- Note the need for two NH₃ in each molecule of urea, so multiply NH₃ by 2.
- 2 NH₃ provides the two H atoms for a molecule of H₂O.
- Each CO₂ provides the O atom for a molecule of H₂O.

(b) synthesis of uranium(VI) fluoride

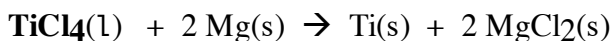
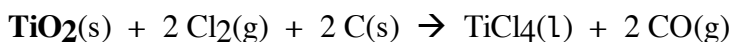


1. The 4 F atoms in UF_4 requires 4 F atoms from HF. (equation 1)

2. The H atoms in HF produce 2 molecules of H_2O . (equation 1)

3. The 1:1 stoichiometry of $\text{UF}_6 : \text{UF}_4$ provides a simple balance. (equation 2)

(c) synthesis of titanium metal from TiO_2 :



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 ($\text{TiO}_2 : \text{TiCl}_4$) 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.

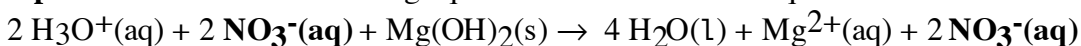
(e) a weak Bronsted acid: the carboxylic acids are weak acids: $\text{CH}_3\text{CO}_2\text{H}$ (acetic)

57. For the following copper salts:

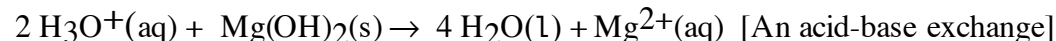
Water soluble: $\text{Cu}(\text{NO}_3)_2$, CuCl_2 — nitrates and chlorides are soluble

Water insoluble: CuCO_3 , $\text{Cu}_3(\text{PO}_4)_2$ — carbonates and phosphates are insoluble

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



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



61. For the reaction of chlorine with NaBr: $\text{Cl}_2(\text{g}) + 2 \text{NaBr}(\text{aq}) \rightarrow 2 \text{NaCl}(\text{aq}) + \text{Br}_2(\text{l})$

(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: **Cl₂** removes the electrons from NaBr

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

63. Reaction: $\text{MgCO}_3(\text{s}) + 2 \text{HCl}(\text{aq}) \rightarrow \text{CO}_2(\text{g}) + \text{MgCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l})$

(a) The net ionic equation: $\text{MgCO}_3(\text{s}) + 2 \text{H}_3\text{O}^+(\text{aq}) \rightarrow \text{CO}_2(\text{g}) + \text{Mg}^{2+}(\text{aq}) + 3 \text{H}_2\text{O}(\text{l})$

The spectator ion is the chloride ion (Cl⁻).

(b) The production of CO₂(g) characterizes this as a gas-forming reaction.

65. Species present in aqueous solutions of:

<u>compound</u>	<u>types of species</u>	<u>species present</u>
(a) NH ₃	molecules (weak base)	NH ₃ , NH ₄ ⁺ , OH ⁻
(b) CH ₃ CO ₂ H	molecules (weak acid)	CH ₃ CO ₂ H, CH ₃ CO ₂ ⁻ , H ⁺
(c) NaOH	ions (strong base)	Na ⁺ and OH ⁻
(d) HBr	ions (strong acid)	H ₃ O ⁺ and Br ⁻

In every case, H₂O will be present (but omitted in this list)

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

(a) $\text{K}_2\text{CO}_3(\text{aq}) + 2 \text{HClO}_4(\text{aq}) \rightarrow 2 \text{KClO}_4(\text{aq}) + \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{l})$ GF

(b) $\text{FeCl}_2(\text{aq}) + (\text{NH}_4)_2\text{S}(\text{aq}) \rightarrow \text{FeS}(\text{s}) + 2 \text{NH}_4\text{Cl}(\text{aq})$ PR

(c) $\text{Fe}(\text{NO}_3)_2(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow \text{FeCO}_3(\text{s}) + 2 \text{NaNO}_3(\text{aq})$ PR

(d) $3 \text{NaOH}(\text{aq}) + \text{FeCl}_3(\text{aq}) \rightarrow 3 \text{NaCl}(\text{aq}) + \text{Fe}(\text{OH})_3(\text{s})$ PR

IN THE LABORATORY

69. For the reaction:

$2 \text{NaI}(\text{s}) + 2 \text{H}_2\text{SO}_4(\text{aq}) + \text{MnO}_2(\text{s}) \rightarrow \text{Na}_2\text{SO}_4(\text{aq}) + \text{MnSO}_4(\text{aq}) + \text{I}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{l})$

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

Reactants: NaI (+1, -1) H₂SO₄ (+1, +6, -2) MnO₂ (+4, -2)

Products: Na₂SO₄ (+1, +6, -2) MnSO₄ (+2, +6, -2) I₂(0) H₂O(+1, -2)

(b) Oxidizing agent: MnO₂ Oxidized: I in NaI

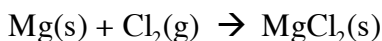
Reducing agent: NaI Reduced: Mn (in MnO₂)

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

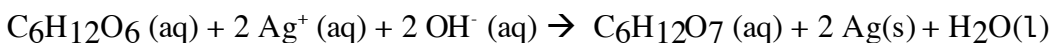
(d) Names of reactants and products:

NaI	H ₂ SO ₄	MnO ₂	Na ₂ SO ₄	MnSO ₄	I ₂	H ₂ O
sodium iodide	sulfuric acid	manganese(IV) oxide	sodium sulfate	manganese(II) sulfate	iodine	water

71. Another way to prepare MgCl₂: Given the reactivity of both elemental magnesium and chlorine, one can bring the two elements into contact (carefully!)



73. In the reaction:



Oxidized: C₆H₁₂O₆ is oxidized to C₆H₁₂O₇ (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: C₆H₁₂O₆ 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).

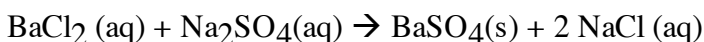
To prove that the establishment of equilibrium is reversible, add strong acid (H₃O⁺). The shift of equilibrium to the left should result in the molecular acid precipitating from solution.

77. Using the reagents: BaCl₂, BaCO₃, Ba(OH)₂, H₂SO₄, Na₂SO₄,

Prepare barium sulfate by:

a precipitation reaction

The reaction of BaCl₂ with Na₂SO₄ will perform this task:



a gas-forming reaction



One might think about using Ba(OH)₂ as one reactant for part (a), but the substance isn't very water-soluble. BaCl₂ is much more water-soluble.