Double Replacement Reactions

A double replacement reaction has the form:

\[ AB + CD \rightarrow AD + CB \]

There are four different possible outcomes to a reaction such as this:

[1] **Formation of a gas.** There are certain compounds which are unstable and decompose to water and a gas. Three common ones are \( \text{H}_2\text{CO}_3 \), \( \text{H}_2\text{SO}_3 \) and \( \text{NH}_4\text{OH} \). They decompose like this:

\[
\text{H}_2\text{CO}_3 \rightarrow \text{H}_2\text{O} + \text{CO}_2 \uparrow \\
\text{H}_2\text{SO}_3 \rightarrow \text{H}_2\text{O} + \text{SO}_2 \uparrow \\
\text{NH}_4\text{OH} \rightarrow \text{H}_2\text{O} + \text{NH}_3 \uparrow
\]

When any of these three compounds appears as a product, write the decomposed form instead.

[2] **Formation of a slightly ionized compound.** Look for compounds like \( \text{H}_2\text{O} \), \( \text{HC}_2\text{H}_3\text{O}_2 \) (acetic acid), \( \text{H}_2\text{C}_2\text{O}_4 \) (oxalic acid) or \( \text{H}_3\text{PO}_4 \) as products. Heat release is the evidence of the formation of these compounds.

[3] **Formation of a precipitate.** Consult the solubility table on page 61–62 in the Chem 061/071 Lab Manual or the User-Friendly Solubility Table from the Learning Centre. “Low solubility” means that very little of the substance dissolves in water, so most of it forms as a precipitate. “Soluble” means that the ions will stay in solution. Just as we use “↑” to indicate a gas has formed, we use “↓” to indicate a precipitate has formed.

[4] **There is no reaction.** None of the above happens, probably because the ions all stay in solution.

**Example 1:** Complete and balance the following equation, if a reaction occurs:

\[ \text{Na}_2\text{CO}_3 + \text{HCl} \rightarrow ? \]

**Solution:**

[1] **Determine what ions are formed.** Consult a list of ions if necessary. The ions in this case are \( \text{Na}^+ \) (not \( \text{Na}_2^+ \)), \( \text{CO}_3^{2-} \), \( \text{H}^+ \), and \( \text{Cl}^- \).

[2] **Form the hypothetical products.** Take into account the valences of the ions involved. The products here would be \( \text{NaCl} \) and \( \text{H}_2\text{CO}_3 \).

[3] **Look for precipitates, slightly ionized compounds and unstable compounds on the product side.** We want to make sure that a reaction will actually occur before we do too much work! In this case, \( \text{NaCl} \) is soluble and so is \( \text{H}_2\text{CO}_3 \), but \( \text{H}_2\text{CO}_3 \) is unstable, so there will be a reaction.

[4] **Write the double replacement equation, if there is a reaction.**

The equation is \( \text{Na}_2\text{CO}_3 + \text{HCl} \rightarrow \text{NaCl} + \text{H}_2\text{CO}_3 \).
Balance the equation, then adjust it for unstable compounds and gases. It’s easier to do it this way than to break down the gases and balance it afterwards.

Na₂CO₃ + 2 HCl → 2 NaCl + H₂O + CO₂↑

Example 2: Complete and balance the following equation, if a reaction occurs:

NaOH + HCl → ?

Solution: We’ll use the same steps as in Example 1.

1. Determine what ions are formed.
   Na⁺, OH⁻, H⁺, Cl⁻.

2. Form the hypothetical products.
   NaCl and H₂O.

3. Look for precipitates, slightly ionized compounds and unstable compounds on the product side.
   NaCl is soluble. H₂O is a slightly ionized compound, so a reaction will occur.

4. Write the double replacement equation, if there is a reaction.
   NaOH + HCl → NaCl + H₂O

5. Balance the equation, then adjust it for unstable compounds and gases.
   It’s balanced as it stands, so we’re done.

Example 3: Complete and balance the following equation, if a reaction occurs:

NaCl + AgNO₃ → ?

Solution: We’ll use the same steps as in Example 1.

1. Determine what ions are formed.
   Na⁺, Cl⁻, Ag⁺, NO₃⁻.

2. Form the hypothetical products.
   NaNO₃ and AgCl.

3. Look for precipitates, slightly ionized compounds and unstable compounds on the product side.
   NaNO₃ is soluble, but AgCl has low solubility, so a reaction will occur.

4. Write the double replacement equation, if there is a reaction.
   NaCl + AgNO₃ → NaNO₃ + AgCl↓

5. Balance the equation, then adjust it for unstable compounds and gases.
   It’s balanced as it stands, so we’re done.
Example 4: Complete and balance the following equation, if a reaction occurs:
NaCl + KNO₃ → ?

Solution: 
[1] **Determine what ions are formed.**
Na⁺, Cl⁻, K⁺, NO₃⁻.

[2] **Form the hypothetical products.**
NaNO₃ and KCl.

[3] **Look for precipitates, slightly ionized compounds and unstable compounds on the product side.**
NaNO₃ and KCl are both soluble, so no reaction will occur. We can stop at this step, since these ions will stay in solution.

EXERCISES
Complete and balance the following equations, if a reaction occurs:

1) BaCl₂ + H₂SO₄ →

2) Na₂CO₃ + HCl →

3) NaC₂H₃O₂ + HCl →

4) K₂CrO₄ + Pb(NO₃)₂ →

5) BiCl₃ + H₂S →

6) SrS + FrClO₃ →

7) K₂C₂O₄ + HCl →

8) H₃PO₄ + Ca(OH)₂ →

9) (NH₄)₂CO₃ + HNO₃ →
10) \[(NH_4)_2CO_3 + CaCl_2 \rightarrow \]

11) \[MgI_2 + Ca(C_2H_3O_2)_2 \rightarrow \]

12) \[KOH + H_3PO_4 \rightarrow \]

13) \[Na_2C_2O_4 + CaCl_2 \rightarrow \]

14) \[(NH_4)_2SO_4 + KOH \rightarrow \]

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**SOLUTIONS**

1) \[BaCl_2 + H_2SO_4 \rightarrow BaSO_4↓ + 2 HCℓ \]

2) \[Na_2CO_3 + 2 HCℓ \rightarrow 2 NaCl + H_2O + CO_2↑ \]

3) \[NaC_2H_3O_2 + HCℓ \rightarrow NaCl + HC_2H_3O_2 \]

4) \[K_2CrO_4 + Pb(NO_3)_2 \rightarrow 2 KNO_3 + PbCrO_4↓ \]

5) \[2 BiCl_3 + 3 H_2S \rightarrow Bi_2S_3↓ + 6 HCℓ \]

6) no reaction

7) \[K_2C_2O_4 + 2 HCℓ \rightarrow 2 KCl + H_2C_2O_4 \]

8) \[2 H_3PO_4 + 3 Ca(OH)_2 \rightarrow 6 H_2O + Ca_3(PO_4)_2↓ \]

9) \[(NH_4)_2CO_3 + 2 HNO_3 \rightarrow 2 NH_4NO_3 + H_2O + CO_2↑ \]

10) \[(NH_4)_2CO_3 + CaCl_2 \rightarrow 2 NH_4Cl + CaCO_3↓ \]

11) no reaction

12) \[3 KOH + H_3PO_4 \rightarrow K_3PO_4 + 3 H_2O \]

13) \[Na_2C_2O_4 + CaCl_2 \rightarrow 2 NaCl + CaC_2O_4↓ \]

14) \[(NH_4)_2SO_4 + 2 KOH \rightarrow 2 NH_3↑ + 2 H_2O + K_2SO_4 \]