Handout *Kimia Dasar I*

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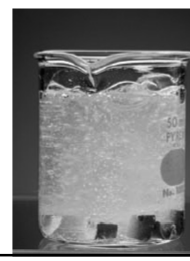
Chemical Reactions






Chemical equation

- All chemical reactions take place according to a set of general principles that relate the amounts of materials consumed in a reaction to the amounts of products formed
  - How much material is needed to make a desired amount of product ?
  - How efficient a chemical synthesis is ?

Chemical equation

- A shorthand expression for a chemical change or reaction.
- Writing chemical equations.

  reactants → products
- Balancing chemical equations
  - Each kind of atom should contains the same number on each side of the equation.
  - The ratio of the number of molecules is equal to the ratio of the number of moles.

- Example:

H_2	+	O_2	→	H_2O
$2H_2$	+	O_2	→	$2H_2O$

2 molecules+ 1 molecule → 2 molecules

2 moles + 1 mole → 2 moles

4.04g + 32.00 g → 36.04 g

36.04 g reactants → 36.04 g products

Sample Question 1 & 2

1. $K_{(s)} + H_2O_{(l)} \rightarrow H_{2(g)} + KOH_{(aq)}$

2. Under appropriate conditions at 1000 °C, ammonia gas reacts with oxygen gas to produce gaseous nitrogen monoxide (common name, nitric oxide) and gaseous water. Write the unbalanced and balanced equations for this reaction.

Group Discussion

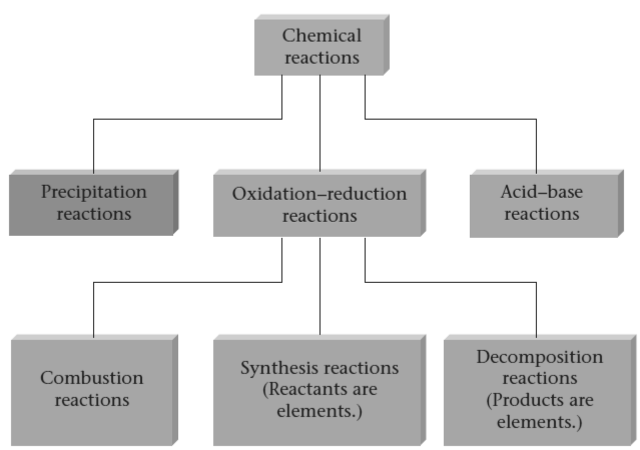
- Balance each of the following chemical equations.

a. $FeCl_{3(aq)} + KOH_{(aq)} \rightarrow Fe(OH)_{3(s)} + KCl_{(aq)}$
b. $Pb(C_2H_3O_2)_{2(aq)} + KI_{(aq)} \rightarrow PbI_{2(s)} + KC_2H_3O_{2(aq)}$
c. $P_4O_{10(s)} + H_2O_{(l)} \rightarrow H_3PO_{4\,(aq)}$
d. $Li_2O_{(s)} + H_2O_{(l)} \rightarrow LiOH_{(aq)}$
e. $MnO_{2(s)} + C_{(s)} \rightarrow Mn_{(s)} + CO_{2(g)}$
f. $Sb_{(s)} + Cl_{2(g)} \rightarrow SbCl_{3(s)}$
g. $CH_{4\,(g)} + H_2O_{(g)} \rightarrow CO_{(g)} + H_{2(g)}$
h. $FeS_{(s)} + HCl_{(aq)} \rightarrow FeCl_{2(aq)} + H_2S_{(g)}$

Reactions

- Why does a chemical reaction occur?
- The most common of the driving forces are:
  1. Formation of a solid
  2. Formation of water
  3. Transfer of electrons
  4. Formation of a gas

Types of reactions



Types of reactions

- **Precipitation reactions**
  - The formation of an insoluble product (precipitate).
  - Usually involve ionic compounds.
  - Ions in solution combine to form a solid salt.

  - Example:
    The precipitation reaction that occurs when yellow potassium chromate, $K_2CrO_{4(aq)}$, is mixed with a colorless barium nitrate solution, $Ba(NO_3)_{2(aq)}$.

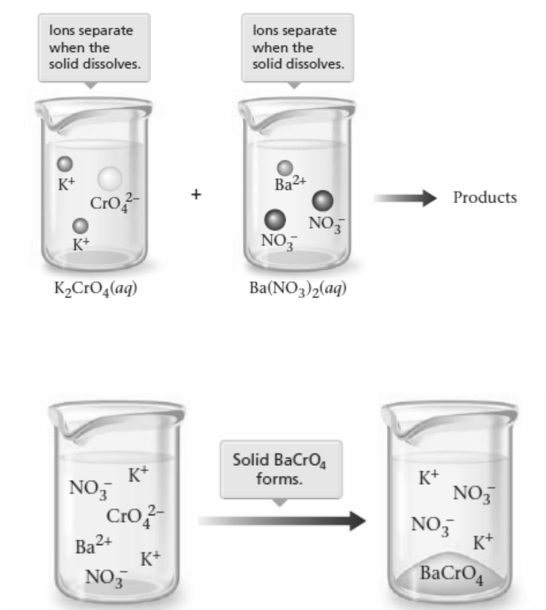
What Happens When an Ionic Compound Dissolves in Water?

The precipitation reaction that occurs when yellow potassium chromate, $K_2CrO_4(aq)$, is mixed with a colorless barium nitrate solution, $Ba(NO_3)_2(aq)$.

- The designation $Ba(NO_3)_{2(aq)}$ means that barium nitrate (a white solid) has been dissolved in water. Note from its formula that barium nitrate contains the Ba^{2+} and NO_3^- ions.
- *In virtually every case when a solid containing ions dissolves in water, the ions separate* and move around independently.
- That is, $Ba(NO_3)_{2(aq)}$ does not contain $Ba(NO_3)_2$ units. Rather, it contains separated Ba^{2+} and NO_3^- ions. In the solution there are two NO_3^- ions for every Ba^{2+} ion.

What Happens When an Ionic Compound Dissolves in Water?

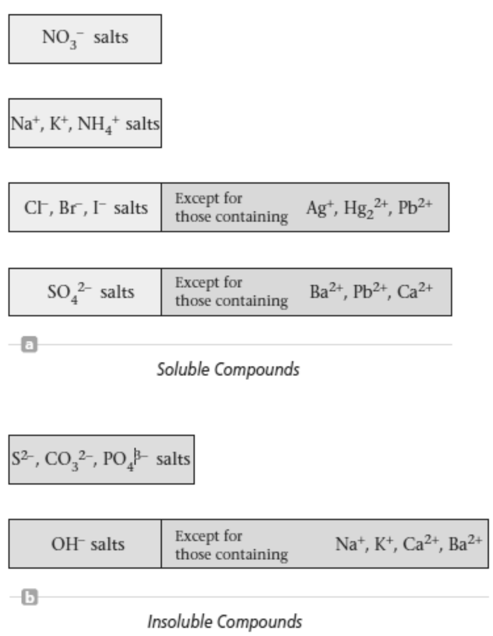
- When each unit of a substance that dissolves in water produces separated ions, the substance is called a **strong electrolyte.**
- Barium nitrate is a strong electrolyte in water, because each $Ba(NO_3)_2$ unit produces the separated ions (Ba_2^+, NO_3^- , NO_3^-).
- Similarly, aqueous K_2CrO_4 also behaves as a strong electrolyte.
- Potassium chromate contains the K^+ and CrO_4^{2-} ions, so an aqueous solution of potassium chromate (which is prepared by dissolving solid K_2CrO_4 in water) contains these separated ions.



General Rules for Solubility of Ionic Compounds (Salts) in Water at 25 °C

1. Most nitrate (NO_3^-) salts are soluble.
2. Most salts of Na^+, K^+, and NH_4^+ are soluble.
3. Most chloride salts are soluble. Notable exceptions are $AgCl$, $PbCl_2$, and Hg_2Cl_2.
4. Most sulfate salts are soluble. Notable exceptions are $BaSO_4$, $PbSO_4$, and $CaSO_4$.
5. Most hydroxide compounds are only slightly soluble.* The important exceptions are NaOH and KOH. $Ba(OH)_2$ and $Ca(OH)_2$ are only moderately soluble.
6. Most sulfide (S^{2-}), carbonate (CO_3^{2-}), and phosphate (PO_4^{3-}) salts are only slightly soluble.*

*The terms *insoluble* and *slightly soluble* really mean the same thing: such a tiny amount dissolves that it is not possible to detect it with the naked eye.



Sample Question 3-5

- Predict what will happen when the following solutions are mixed. Write the balanced equation for any reaction that occurs.

3. $KNO_{3(aq)}$ and $BaCl_{2(aq)}$
4. $Na_2SO_{4(aq)}$ and $Pb(NO_3)_{2(aq)}$
5. $KOH_{(aq)}$ and $Fe(NO_3)_{3(aq)}$

Complete Net Ionic Equation

$$\mathbf{Ba(NO_3)_{2(aq)} + K_2CrO_{4(aq)} \rightarrow BaCrO_{4(s)} + 2KNO_{3(aq)}}$$

- This is called the **molecular equation** for the reaction; it shows the complete formulas of all reactants and products.
- However, although this equation shows the reactants and products of the reaction, it does not give a very clear picture of what actually occurs in solution.
- The **complete ionic equation,** better represents the actual forms of the reactants and products in solution.

$$\underbrace{2K^{+}(aq) + CrO_4^{2-}(aq)}_{\text{Ions from } K_2CrO_4} + \underbrace{Ba^{2+}(aq) + 2NO_3^{-}(aq)}_{\text{Ions from } Ba(NO_3)_2} \rightarrow BaCrO_4(s) + 2K^{+}(aq) + 2NO_3^{-}(aq)$$

- The complete ionic equation reveals that only some of the ions participate in the reaction.
- K^+ and NO_3^- ions are present in solution both before and after the reaction. Ions such as these, which do not participate directly in a reaction in solution, are called **spectator ions.**
- The ions that participate in this reaction are the Ba^{2+} and CrO_4^{2-} ions, which combine to form solid BaCrO4:

$$Ba^{2+}(aq) + CrO_4^{2-}(aq) \rightarrow BaCrO_4(s)$$

- **Acid - base reactions**

- One of the most fundamental chemical reactions is the combination of a hydroxide ion and a hydronium ion to produce two molecules of water:

  $OH^-_{(aq)} \quad + \quad H_3O^+_{(aq)} \quad \rightarrow \quad 2H_2O_{(l)}$
- Acid as a proton donor and base as a proton acceptor.
- Any reaction in which H^+ moves from one species to another.
- Also called a neutralization reaction.

acid + base → salt + water

- **Strong acids:**
  React with water to produce hydronium ions.
  Ex: HNO_3, $HClO_4$, H_2SO_4, HCl, HBr and HI.
- **Weak acids**
  Substances that can't readily donate protons to water molecule.
  Ex: H_3PO_4, CH_3COOH
- **Strong base**
  Substances that are completely ionized into the metals ions and hydroxide ions.
  Ex: NaOH, $Ba(OH)_2$
- **Weak base**
  $Al(OH)_3$, NH_3

Sample Question 6

6. Nitric acid is a strong acid. Write the molecular, complete ionic, and net ionic equations for the reaction of aqueous nitric acid and aqueous potassium hydroxide.

Besides water, which is *always a product* of the reaction of an acid with OH, the second product is an ionic compound, which might precipitate or remain dissolved, depending on its solubility.

$$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$$
$$HNO_3(aq) + KOH(aq) \rightarrow H_2O(l) + KNO_3(aq)$$

Dissolved ionic compounds

This ionic compound is called a **salt.**

- **Base and acid oxides**

\- Metals-oxides are bases. The oxide anion, O_2^-, is a strong base that readily accepts a proton from a water molecule.

$Na_2O_{(s)}$ + $H_2O_{(l)}$ → $2Na^+_{(aq)}$ + $2OH^-_{(aq)}$

$CaO_{(s)}$ + $H_2O_{(l)}$ → $Ca(OH)_{2(s)}$

- Non metal oxides react with the water to produce acids. Non metal oxides whose molecules contain one non metal atom react in a 1:1 ratio with water:

$SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)}$
$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$
$CO_{2(g)} + H_2O_{(l)} \rightarrow H_2CO_{3(aq)}$

- Oxide whose molecules contain more than one non metal atom react to form more than one molecule of acid:

$N_2O_{5(g)} \quad + \quad H_2O_{(l)} \quad \rightarrow \quad 2HNO_{3(aq)}$
$Cl_3O_{7(g)} \quad + \quad H_2O_{(l)} \quad \rightarrow \quad 2HClO_{4(aq)}$
$P_4O_{10(s)} \quad + \quad 6H_2O_{(l)} \quad \rightarrow \quad 4H_3PO_{4(aq)}$

- **Oxidation – reduction reactions**
- Oxidation – reduction reactions occur when electrons from one chemical substance are transferred to another.
- Example:

$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$

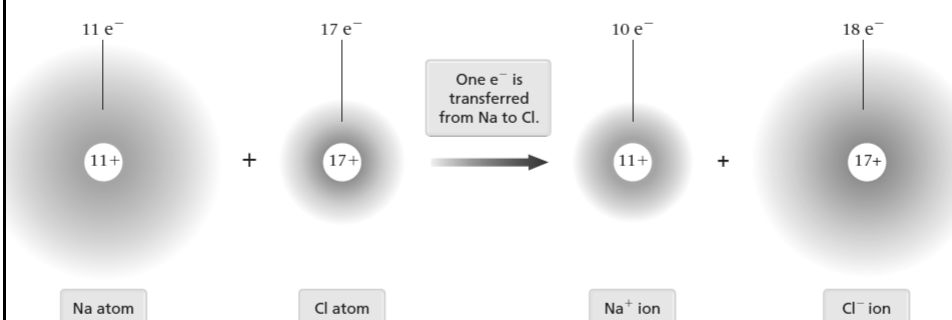
$2Mg_{(s)} + O2_{(g)} \rightarrow 2MgO_{(s)}$

$2Al_{(s)} + Fe_2O_{3(s)} \rightarrow 2Fe_{(s)} + Al_2O_{3(s)}$

- Sodium metal is composed of sodium atoms, each of which has a net charge of zero. (The positive charges of the 11 protons in its nucleus are exactly balanced by the negative charges on the 11 electrons.) Similarly, the chlorine molecule consists of 2 uncharged chlorine atoms (each has 17 protons and 17 electrons).
- However, in the product (sodium chloride), the sodium is present as Na and the chlorine as Cl. By what process do the neutral atoms become ions?

$$\underbrace{\mathrm{Na} + \mathrm{Cl}}_{e^-} \rightarrow \mathrm{Na}^+ + \mathrm{Cl}^-$$

- After the electron transfer, each sodium has ten electrons and eleven protons (a net charge of 1), and each chlorine has eighteen electrons and seventeen protons (a net charge of 1).



- **Metal displacement**

- The reaction occurs when one metal in solution is displaced by another metal by means of a redox reaction.
- Oxidation: $Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$
- Reduction: $Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu_{(s)}$
- Redox: $Zn_{(s)} + Cu^{2+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + Cu_{(s)}$

- **Oxidation of metals by H_3O^+ and H_2O**

- $2Li_{(s)} + 2H_3O^+ \rightarrow 2Li^+_{(aq)} + H_{2(g)} + 2H_2O_{(l)}$
- $2Fe_{(s)} + 4H_3O^+ \rightarrow 2Fe^{2+}_{(aq)} + 2H_{2(g)} + 4H_2O_{(l)}$
- $2Al_{(s)} + 6H_3O^+ \rightarrow 2Al^{3+}_{(aq)} + 3H_{2(g)} + 6H_2O_{(l)}$

- $2Na \rightarrow 2Na^+ + 2e^-$
- $\underline{2H_2O + 2e^- \rightarrow H_2 + 2OH^-}$

$2Na + 2H_2O \rightarrow H_2 + 2Na^+ + 2OH^-$

- **Oxidation by molecular oxygen**

- Almost all elements combine with molecular oxygen to form binary oxides; the loss of electrons is called oxidation because elements lose electrons when they combine with oxygen. The more easily a metal is oxidized, the more readily it reacts with molecular oxygen.

The stoichiometry of chemical reactions

- The knowledge on the relationship among atoms, moles, and masses including how much they are present in combination with the concept of a balanced chemical equation is required.

Sample Question 7

7. How many grams of hydrogen do we need to produce 68 g of ammonia?

   $N_2 + 3H_2 \rightarrow 2NH_3$

Sample Question 8

8. Poisonous hydrogen cyanide (HCN) is an important industrial chemical. It is produced from methane , ammonia, and molecular oxygen. The reaction also produces water. An industrial manufacturer wants to convert 175 kg of methane into HCN. How much molecular oxygen will be required for this synthesis?

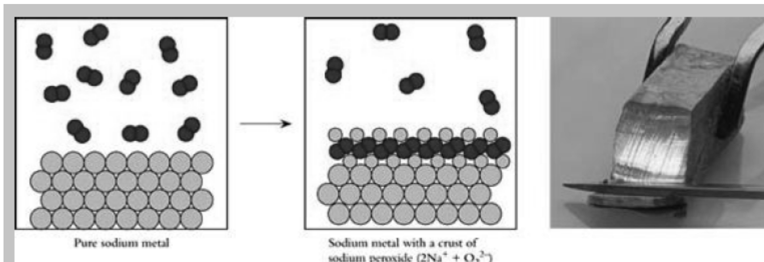
$CH_4 + NH_3 + O_2 \rightarrow HCN + H_2O$

Yields of chemical reactions

- The amount of a product obtained from a reaction is often described in terms of the yield of the reaction.
- The quantity of product predicted by stoichiometry → the theoretical yield
- the amount actually obtained → the actual yield

Percent yield = (actual yield) / (theoretical yield) (100%)

- Under practical conditions, chemical reactions almost always produce smaller amounts of products than the amounts predicted by stoichiometric analysis.
- There are three major reasons for this:
  – Many reactions stop before reaching completion. Other reactions do not go to completion because they reach dynamic equilibrium. While reactant molecules continue to form product molecules, product molecules also interact to re-form reactant molecules.



The reaction of sodium metal with O_2 gas produces a crust of sodium peroxide, which blocks additional reactant molecules from reaching the unreacted metal.

– Competing reactions often consume some of the starting materials.
– When the product of a reaction is purified and isolated, some of it is inevitably lost during the collection process. Gases may escape while being pumped out of a reactor. Liquids adhere to glass surfaces, making it impossible to transfer every drop of a liquid product. Likewise, it is impossible to scrape every trace of a solid material from a reaction vessel.

Sample Question 9

- The Haber synthesis of ammonia stops when 13% of the starting materials have formed products. Knowing this, how much ammonia could an industrial producer expect to make from 2.0 metric tons of molecular hydrogen?
- $N_2 + 3H_2 \rightarrow 2NH_3$

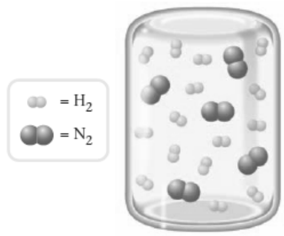
Sample Question 10

- The industrial production of hydrogen cyanide is described in Example . If the yield of this synthesis is 97.5%, how many kilograms of methane should be used to produce 1.5×10^5 kg of HCN?

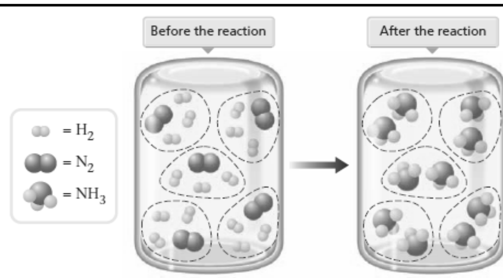
The limiting reagent

- often chemical reactions are run with an excess of one or more starting materials
- One reactant will “run out” before the others.
- The reactant that runs out is called the limiting reagent because it limits how much product can be made.
- The other starting materials are said to be in excess.

- $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$
- Consider the following container of $N_{2(g)}$ and $H_{2(g)}$:



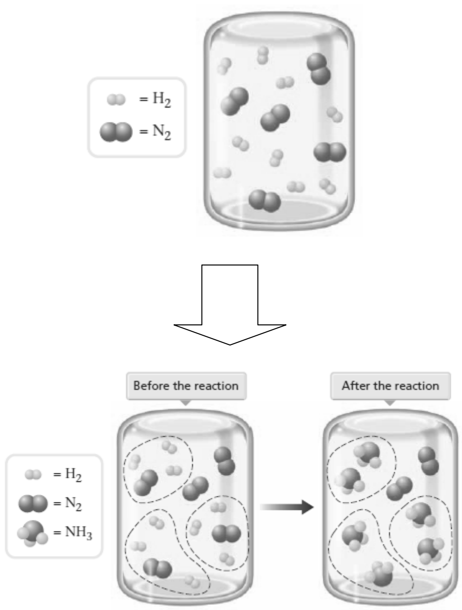
- Each N_2 requires $3H_2$ molecules to form $2NH_3$



- In this case, the mixture of N_2 and H_2 contained just the number of molecules needed to form NH_3 with nothing left over. That is, the ratio of the number of H_2 molecules to N_2 molecules was $15H_2 : 5N_2 = 3H_2 : 1H_2$
- This ratio exactly matches the numbers in the balanced equation:

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

- This type of mixture is called a *stoichiometric mixture.*



- In this case H_2 is limiting. That is, the H_2 molecules are used up before all of the N_2 molecules are consumed.
- In this situation, the amount of hydrogen limits the amount of product (ammonia) that can form—hydrogen is the limiting reactant.
- Some N_2 molecules are left over in this case because the reaction runs out of H_2 molecules first.

The reactant that runs out first and thus limits the amounts of products that can form is called the **limiting reactant (limiting reagent).**

Steps for Solving Stoichiometry Problems Involving Limiting Reactants

Step 1 Write and balance the equation for the reaction.

Step 2 Convert known masses of reactants to moles.

Step 3 Using the numbers of moles of reactants and the appropriate mole ratios, determine which reactant is limiting.

Step 4 Using the amount of the limiting reactant and the appropriate mole ratios, compute the number of moles of the desired product.

Step 5 Convert from moles of product to grams of product, using the molar mass (if this is required by the problem).

Sample Question 11

- For the ammonia synthesis, if we start with 84.0 g of molecular nitrogen and 24.2 g of molecular hydrogen, what mass of ammonia can be prepared?
- $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$

Sample Question 12

- Nitric acid is used in the production of fertilizers and explosives. One step in the industrial production of nitric acid is the reaction of ammonia with molecular oxygen to form nitrogen oxide:

  $4\ NH_3 + 5\ O_2 \rightarrow 4\ NO + 6\ H_2O$

  In a study of this reaction, a chemist mixed 125 g of ammonia with 256 g of oxygen and allowed them to react to completion. What masses of NO and were produced, and what mass of which reactant was left over?