Chemical Reactions

Chapter 4

Chemical Equations

A balanced chemical equation gives us a variety of information about the <u>relative</u> amounts of substances reacting and forming.

- Law of Conservation of Mass - # of atoms of given element on reactant side must equal # of atoms of that element on product side.
- coefficients of balanced equation tell *how many* of each species are involved in reaction
- when balancing equation - *identities* (subscripts) cannot be changed; only change *amounts* (coefficients)

Patterns of Chemical Reactions

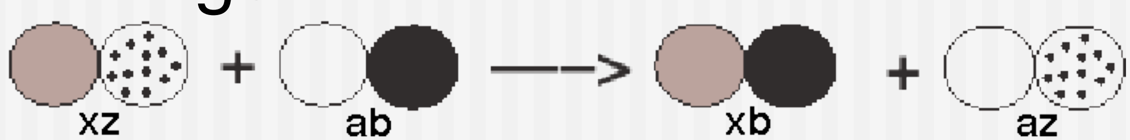
Four general types of chemical reactions

- combination

- decomposition

- displacement

- exchange



Combination Reactions

- often occur with O or halogens
  - metal + oxygen → metal oxide

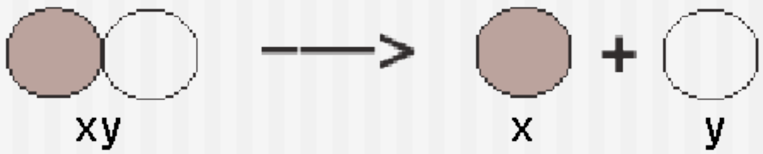
    $4\ Fe(s)\ +\ 3\ O_2(g) \rightarrow 2\ Fe_2O_3\ (s)$
  - metal + halogen → ionic compound

    $2\ K(s)\ +\ F_2(g) \rightarrow 2\ KF\ (s)$

    $2\ Fe(s)\ +\ 3\ C_2(g) \rightarrow 2\ FeCl_3(s)$
  - nonmetal + oxygen (or halogen) → molecular compound

    $P_4(s)\ +\ 5\ O_2(g) \rightarrow 2\ P_2O_5(s)$

Decomposition Reactions



- thermal decomposition
  - metal oxides

    $2\ HgO(s) \rightarrow 2\ Hg(l) + O_2(g)$

  - metal carbonates

    $MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$

- unstable compounds

  $4\ C_3H_5(NO_3)_3(l) \rightarrow 12\ CO_2(g) + 10\ H_2O(g) + 6\ N_2(g) + O_2(g)$

Displacement Reactions

examples:

$$3\ C(s)\ +\ Fe_2O_3(g) \rightarrow 3\ CO(g)\ +\ 2\ Fe(l)$$

$$3\ Fe(s)\ +\ 4\ H_2O(g) \rightarrow Fe_3O_4(s) + 4\ H_2(g)$$

Exchange Reactions

most often seen in precipitation reactions

$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

Steps for Balancing Chemical Equations

1. write an unbalanced equation with correct reactant and product formulas

2. balance atoms of one element (start with non-H or O if possible)

3. balance atoms of remaining elements

4. check equation to verify that it is balanced

Example 1

Iron metal reacts with chlorine gas to form iron(III) chloride. Write a balanced equation for this reaction.

1. $Fe(s) + Cl_2(g) \rightarrow FeCl_3(s)$
2. $Fe(s) + 3\ Cl_2(g) \rightarrow 2\ FeCl_3(s)$
3. $2\ Fe(s) + 3\ Cl_2(g) \rightarrow 2\ FeCl_3(s)$
4. 2 Fe, 6 Cl $\rightarrow$ 2 Fe, 6 Cl

Example 2

Write and balance the chemical equation for the combustion of $C_7H_{14}O_2$.

$$2\ C_7H_{14}O_2 + 19\ O_2 \rightarrow 14\ CO_2 + 14\ H_2O$$

$$14\ C,\ 28\ H,\ 42\ O \rightarrow 14\ C,\ 28\ H,\ 42\ O$$

Example 3

Balance the following chemical equation:

$NaBr + H_3PO_4 \rightarrow HBr + Na_3PO_4$

$3\ NaBr + H_3PO_4 \rightarrow 3\ HBr + Na_3PO_4$

Moles and Chemical Reactions

$$N_2 + 3\,H_2 \rightarrow 2\,NH_3$$

1 molecule N_2	3 molecules H_2	2 molecules NH_3
1 mol N_2	3 mol H_2	2 mol NH_3
28 g N_2	6 g H_2	34 g NH_3

34 g total (28 g N_2 + 6 g H_2)

34 g total (34 g NH_3)

Example 4

How many moles of ammonia are formed from the reaction of 0.30 mol nitrogen gas with excess hydrogen gas?

$$N_2 + 3H_2 \rightarrow 2NH_3$$

$$\textbf{mol NH}_3 = (0.30 \cancel{\text{mol}} \text{ N}_2)\left(\frac{2 \text{ mol NH}_3}{1 \cancel{\text{mol}} \text{ N}_2}\right) = 0.60 \text{ mol NH}_3$$

Example 5

Propane, C_3H_8, is a common fuel used for cooking and home heating. What mass of O_2 is consumed in the combustion of 1.00 g of propane?

$$2\ C_3H_8 + 10\ O_2 \rightarrow 6\ CO_2 + 8\ H_2O$$

$$\textbf{mol } C_3H_8 = (1.00\text{ g})\left(\frac{1\text{ mol}}{44.0962\text{ g}}\right) = 0.0227\text{ mol } C_3H_8$$

$$\textbf{mol } O_2 = (0.0227\text{ mol } C_3H_8)\left(\frac{10\text{ mol } O_2}{2\text{ mol } C_3H_8}\right) = 0.113\text{ mol } O_2$$

$$\textbf{mass } O_2 = (0.113\text{ mol})\left(\frac{31.9899\text{ g}}{1\text{ mol}}\right) = 3.63\text{ g}$$

Example 6

When heated in dry air, sodium reacts with oxygen to form sodium peroxide: $Na(s) + O_2(g) \rightarrow Na_2O_2(s)$

What masses (in kilograms) of Na and O_2 are required to prepare 457 kg of Na_2O_2 by this reaction?

$$2\ Na(s) + O_2(g) \rightarrow Na_2O_2(s)$$

$$\text{kmol } Na_2O_2 = (457\text{ kg})\left(\frac{1\text{ kmol}}{77.9786\text{ kg}}\right) = 5.86\text{ kmol}$$

$$\text{mass of Na} = (5.86\text{ kmol } Na_2O_2)\left(\frac{2\text{ kmol Na}}{1\text{ kmol } Na_2O_2}\right)\left(\frac{22.9899\text{ kg Na}}{1\text{ kmol}}\right) = 269\text{ kg Na}$$

$$\text{mass of } O_2 = (5.86\text{ kmol } Na_2O_2)\left(\frac{1\text{ kmol } O_2}{1\text{ kmol } Na_2O_2}\right)\left(\frac{31.9988\text{ kg } O_2}{1\text{ kmol}}\right) = 188\text{ kg } O_2$$

Limiting Reactants

The reactant that is completely comsumed in a chemical reaction is called the limiting reactant.

$$2\ H_2(g)\ +\ O_2(g) \rightarrow 2\ H_2O\ (l)$$

If you start with 10 mol H_2 and 7 mol O_2, which reactant is used up first?

Try H_2:

$$\textbf{mol } H_2O \textbf{ formed} = (10 \text{ mol } H_2)\left(\frac{2 \text{ mol } H_2O}{2 \text{ mol } H_2}\right) = 10 \text{ mol } H_2O$$

Try O_2:

$$\textbf{mol } H_2O \textbf{ formed} = (7 \text{ mol } O_2)\left(\frac{2 \text{ mol } H_2O}{1 \text{ mol } O_2}\right) = 14 \text{ mol } H_2O$$

smallest amount formed:
H_2 is limiting reactant

Example 7

Suppose 184 g of Al_2O_3, 75 g of C, and 92 g of Cl_2 are mixed and the reaction proceeds until one of the reactants is used up. Which is the limiting reactant? How many grams of $AlCl_3$ are produced, and how many grams of the other reactants remain?

$$2\ Al_2O_3 \ + \ 3\ C \ + \ 6\ Cl_2 \ \rightarrow \ 4\ AlCl_3 \ + \ 3\ CO_2$$

$$\underset{\text{from } Al_2O_3}{g\ AlCl_3} = (184\ g)\left(\frac{1\ mol\ Al_2O_3}{101.9612\ g}\right)\left(\frac{4\ mol\ AlCl_3}{2\ mol\ Al_2O_3}\right)\left(\frac{133.3405\ g}{1\ mol\ AlCl_3}\right) = 481\ g\ AlCl_3$$

$$\underset{\text{from } C}{g\ AlCl_3} = (75\ g)\left(\frac{1\ mol\ C}{12.011\ g}\right)\left(\frac{4\ mol\ AlCl_3}{3\ mol\ C}\right)\left(\frac{133.3405\ g}{1\ mol\ AlCl_3}\right) = 1110\ g\ AlCl_3$$

$$\underset{\text{from } Cl_2}{g\ AlCl_3} = (92\ g)\left(\frac{1\ mol\ Cl_2}{70.906\ g}\right)\left(\frac{4\ mol\ AlCl_3}{6\ mol\ Cl_2}\right)\left(\frac{133.3405\ g}{1\ mol\ AlCl_3}\right) = 115\ g\ AlCl_3$$

Cl_2 limiting reactant

mass of $AlCl_3$ produced

Example 7 (cont'd)

mass of Al_2O_3 remaining:

$$\text{mass } Al_2O_3 \text{ used} = (92 \text{ g } Cl_2)\left(\frac{1 \text{ mol } Cl_2}{70.906 \text{ g}}\right)\left(\frac{2 \text{ mol } Al_2O_3}{6 \text{ mol } Cl_2}\right)\left(\frac{101.9612 \text{ g}}{1 \text{ mol}}\right) = 44 \text{ g } Al_2O_3$$

$$\text{mass } Al_2O_3 \text{ remaining} = 184 \text{ g} - 44 \text{ g} = 140. \text{ g}$$

mass of C remaining:

$$\text{mass C used} = (92 \text{ g } Cl_2)\left(\frac{1 \text{ mol } Cl_2}{70.906 \text{ g}}\right)\left(\frac{3 \text{ mol C}}{6 \text{ mol } Cl_2}\right)\left(\frac{12.011 \text{ g}}{1 \text{ mol}}\right) = 7.8 \text{ g C}$$

$$\text{mass C remaining} = 75 \text{ g} - 7.8 \text{ g} = 67 \text{ g}$$

Percent Yield

Amount of product calculated to form when all the limiting reactant is consumed is the *theoretical yield*.

Actual yield is the amount actually obtained from the reaction.

$$\% \text{ yield} = \frac{\textbf{actual yield}}{\textbf{theoretical yield}} \textbf{ x 100\%}$$

Example 8

Calculate the percent yield of C_2H_5Cl if the reaction of 125 g of C_2H_6 with 255 g of Cl_2 produced 206 g of C_2H_5Cl.

$$C_2H_6 + Cl_2 \rightarrow C_2H_5Cl + HCl$$

determine limiting reactant:

$$\text{mass of } C_2H_5Cl \text{ from } C_2H_6 = (125\text{ g})\left(\frac{1\text{ mol } C_2H_6}{30.0694\text{ g}}\right)\left(\frac{1\text{ mol } C_2H_5Cl}{1\text{ mol } C_2H_6}\right)\left(\frac{64.5145\text{ g}}{1\text{ mol}}\right) = 268\text{ g}$$

$$\text{mass of } C_2H_5Cl \text{ from } Cl_2 = (255\text{ g})\left(\frac{1\text{ mol } Cl_2}{70.906\text{ g}}\right)\left(\frac{1\text{ mol } C_2H_5Cl}{1\text{ mol } Cl_2H_6}\right)\left(\frac{64.5145\text{ g}}{1\text{ mol}}\right) = 232\text{ g}$$

Cl_2 limiting reactant

theoretical yield of C_2H_5Cl

% yield = (206 g / 232 g) x 100 % = 88.8 %

Combustion Analysis

During combustion, a compound is burned in oxygen. The carbon in the compound is converted to CO_2 and the hydrogen is converted to H_2O. Many compounds also contain oxygen.

Mass of oxygen = mass of sample - (mass of C + mass of H)

Combustion analysis will tell us how much CO_2 and H_2O were formed during the reaction. This information can be used to determine the amounts of C, H and O in the sample. From this we can get the empirical formula.

Example 9

Propionic acid, an organic acid, contains only C, H, and O. If 0.236 g of the acid burns completely in O_2 and gives 0.421 g of CO_2 and 0.172 g of H_2O, what is the empirical formula of the acid?

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$$

determine masses of C and H in original sample:

$$\text{mass of C} = (0.421 \text{ g } CO_2)\left(\frac{1 \text{ mol } CO_2}{44.0098 \text{ g}}\right)\left(\frac{1 \text{ mol C}}{1 \text{ mol } CO_2}\right)\left(\frac{12.011 \text{ g}}{1 \text{ mol C}}\right) = 0.115 \text{ g C}$$

$$\text{mass of H} = (0.172 \text{ g } H_2O)\left(\frac{1 \text{ mol } H_2O}{18.0152 \text{ g}}\right)\left(\frac{2 \text{ mol H}}{1 \text{ mol } H_2O}\right)\left(\frac{1.0079 \text{ g}}{1 \text{ mol H}}\right) = 0.0192 \text{ g H}$$

mass of O in compound = 0.236 g - (0.115 g + 0.0192 g) = 0.102 g O

determine masses of C and H in original sample:

Example 9 (cont'd)

Propionic acid, an organic acid, contains only C, H, and O. If 0.236 g of the acid burns completely in O_2 and gives 0.421 g of CO_2 and 0.172 g of H_2O, what is the empirical formula of the acid?

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$$

determine mole ratios and empirical formula as usual:

$$\left.\begin{array}{l} \text{mol of C} = (0.115\text{ g C})\left(\dfrac{1\text{ mol C}}{12.011\text{ g}}\right) = 0.00957\text{ mol C} \rightarrow 1.5011 \rightarrow \mathbf{3} \\ \text{mol of H} = (0.0192\text{ g H})\left(\dfrac{1\text{ mol H}}{1.0079\text{ g}}\right) = 0.0191\text{ mol H} \rightarrow 2.994 \rightarrow \mathbf{6} \\ \text{mol of O} = (0.102\text{ g O})\left(\dfrac{1\text{ mol O}}{15.9994\text{ g}}\right) = 0.00638\text{ mol O} \rightarrow 1.000 \rightarrow \mathbf{2} \end{array}\right\} C_3H_6O_2$$

divide by smallest