CHEMICAL EQUATIONS

Parts of a Chemical Equation

$$4\ Al\ (s) + 3\ O_2\ (g) \longrightarrow 2\ Al_2O_3\ (s)$$

Reactants (on the left)

Products (on the right)

read as "produces" or "yields"

The plus sign

is read as "plus" or "and"

The numbers 4, 3 and 2 that are before chemical formulae are called coefficients.

(s), (g), and (l) are the physical states of the compounds taking part in the chemical reaction.

Symbols Used in Chemical Equations

- **Solid (s)**
- Light $\xrightarrow{h\nu}$
- **Liquid (l)**
- **Gas (g)**
- **Aqueous solution (aq)**
- **Catalyst:** $\xrightarrow{\text{name of catalyst}}$
- **Escaping gas (↑)**
- **Change of temperature (Δ)**

Balancing of Chemical Equations

Because of the principle of the **conservation of matter**, an **equation must be balanced,**

i.e. an equation must have the same number of atoms on both sides.

Law of conservation of mass

Mass is neither created or destroyed in a chemical reaction

How to Balance a Chemical Equation

If the number of atoms of each element on each side of the chemical equation are **not** the same then **whole numbers** called **COEFFICIENTS** are used so that the number of atoms of each element on each side of the chemical equation are the same.

You may <u>not</u> change the subscripts in the formulas of reactants or products taking part in the chemical reaction.

Steps in Balancing a Chemical Equation

Example: Balance the chemical equation:

$$Al\ (s) \quad + \quad O_2\ (g) \longrightarrow Al_2O_3\ (s)$$

1. List the number of each atom of the elements that are products in a column right next to the reactants; atoms that are the same side by side.

Reactants	**Products**
Al	2 Al
2 O	3 O

The **subscript tells us the number of atoms of that element.**

Steps in Balancing a Chemical Equation

$$Al\ (s) + O_2\ (g) \longrightarrow Al_2O_3\ (s)$$

2. Multiply the number of atoms by small whole numbers so that the same number of atoms on both sides is obtained.

Reactants	Products
2 x Al	2 Al
3 x 2 O = 6 O	2 x 3 O = 6 O

3. Write the equation again with the numbers **that you multiplied the reactants and products by** as coefficients.

$$2Al\ (s) + 3O_2\ (g) \longrightarrow 2Al_2O_3\ (s)$$

Notice coefficient is placed in front of formula containing O

$$2Al\,(s) + 3O_2\,(g) \longrightarrow 2Al_2O_3\,(s)$$

4. Check the equation again to see if it is balanced.

Reactants	**Products**
2 Al	4 Al Still not balanced
3 x 2 O = 6 O	2 x 3 O = 6 O

5. Proceed once again to multiply the chemical equation by small whole numbers so that the same number of atoms on both sides is obtained.

Reactants	**Products**
2 x 2 Al = 4 Al	4 Al
3 x 2 O = 6 O	2 x 3 O = 6 O

6. Write the equation again with the numbers that you multiplied by as coefficients.

$$4Al\ (s)\ +\ 3O_2\ (g) \longrightarrow 2Al_2O_3\ (s)$$

Number of atoms on both sides of the equation are now equal.

Equation is now balanced!

Pointers for Balancing Chemical Equations

- **Before balancing an equation, check that the chemical formulas of ALL substances are correct.**

- Do not write the number 1 as a coefficient. The number 1 is assumed and does not appear in the balanced chemical equation.

- Begin balancing the equation starting with the most complex formula.

- Take one element at a time, working left to right except for H and O. Save H for next to last, and O until last.

Pointers for Balancing Chemical Equations

- Occasionally, it is helpful to use a fractional coefficient to balance an element in a diatomic molecule; for example,

$$H_2\ (g) + \frac{1}{2} O_2\ (g) \longrightarrow H_2O\ (l)$$

Multiplying entire equation above by 2 :

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

- Treat polyatomic ions as single units if they **DO** remain intact during the chemical reaction. If they **do not** you must treat them as individual atoms when balancing chemical equations.

What is meant by a polyatomic ion remaining intact?

This means that the polyatomic ion observed in formula of reactant(s) is observed as part of the formula of product(s).

Example: $Pb(NO_3)_2\ (aq) + KI\ (aq) \longrightarrow PbI_2\ (s) + KNO_3\ (aq)$

Polyatomic ion is intact on both sides of unbalanced equation

List polyatomic ion as a unit:

Reactants	**Products**
1 Pb	1 Pb
2 NO_3	1 NO_3
1 K	1 K
1 I	2 I

Example:

$$Ca(NO_3)_2 \longrightarrow Ca(NO_2)_2 + O_2$$

The polyatomic ion, NO_3^- has not remained intact, i.e. it does **not** appear on the right hand side of the equation, so list all atoms when attempting to balance this chemical equation

Reactants	**Products**
1 Ca	1 Ca
2 N	2 N
6 O	6 O

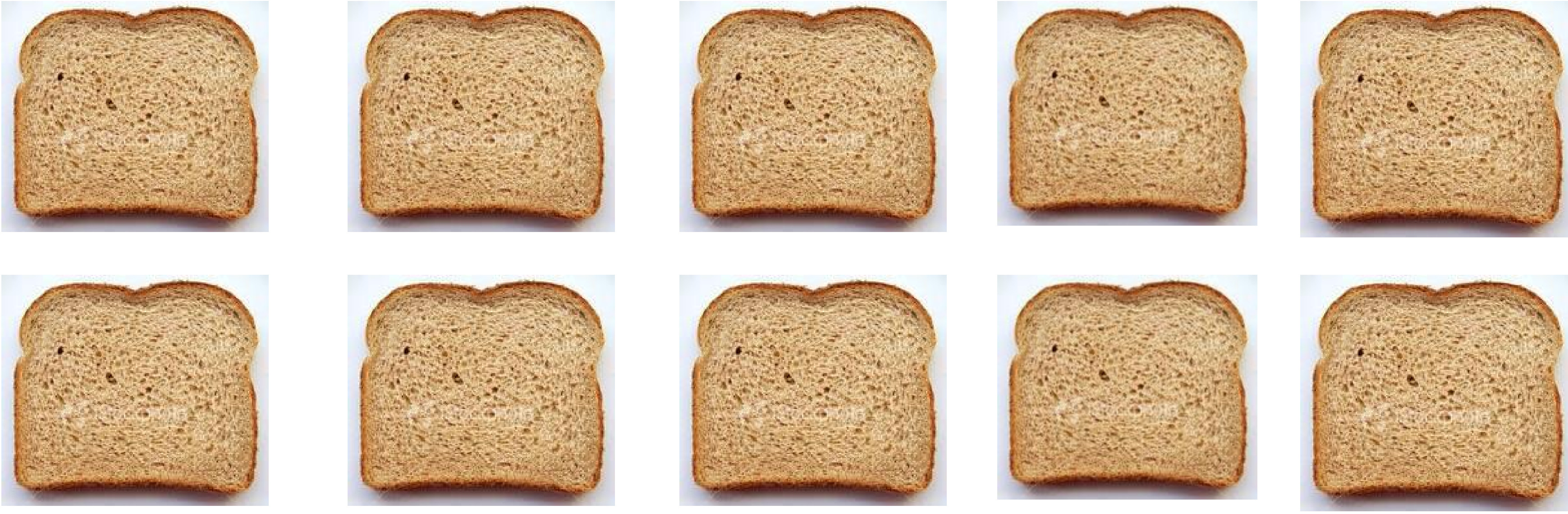
This chemical equation is already balanced

Limiting Reagent

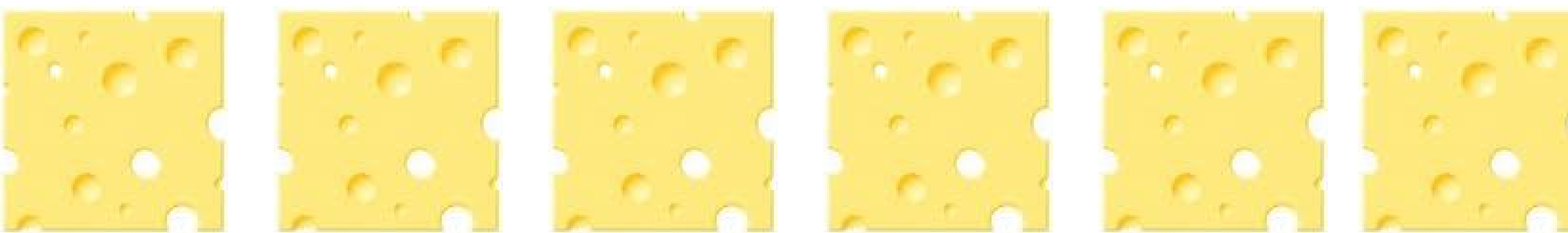
- The limiting reagent is the reactant in a chemical reaction which limits the amount of products that can be formed.

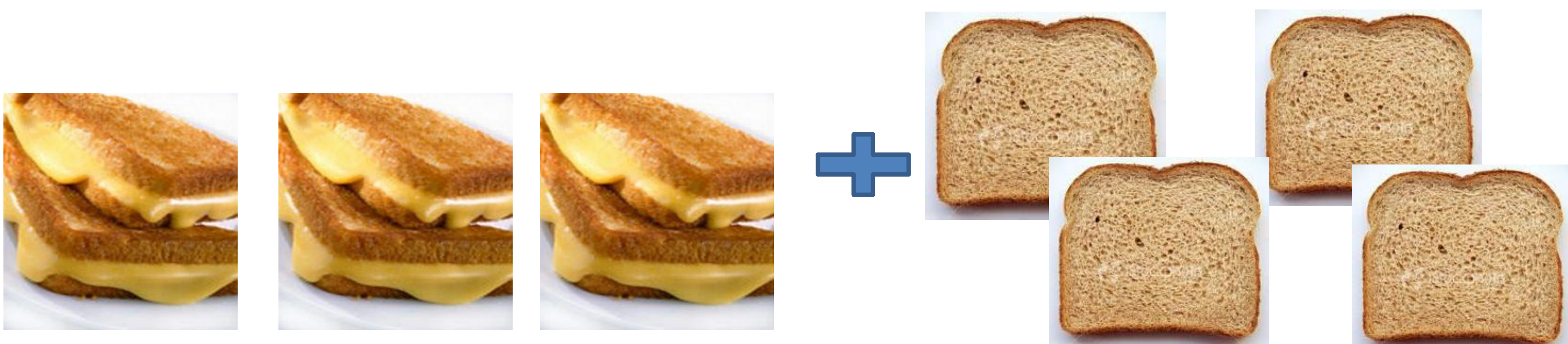
- The limiting reagent in a chemical reaction is present in **insufficient** quantity to consume the other reactant(s).

- This situation arises when reactants are mixed in **non-stoichiometric** ratios.



Cheese is the limiting reagent





Stoichiometric Ratio

$$2H_{2\,(g)} + O_{2\,(g)} \rightarrow 2H_2O_{(l)}$$

atomic	4 H atoms	2 O atoms	4 H atoms 2 O atoms
molecular	2 H_2	1 O_2	2 H_2O
molar	2 mol H_2	1 mol O_2	2 mol H_2O
mass	2 x 2.016 g H_2	32.00 g O_2	2 x 18.02 g H_2O

$$\underbrace{2 \times 2.016\text{ g } H_2 \quad 32.00\text{ g } O_2}_{36.04\text{ g reactants}} \qquad 2 \times 18.02\text{ g } H_2O$$

36.04 g products

36.04 g reactants

Using Stoichiometric Ratios

$$2\,NH_3\,(g) + 5/2\,O_2\,(g) \longrightarrow 2NO\,(g) + 3\,H_2O\,(l)$$

Fraction is removed by multiplying **all coefficients** by 2

$$4NH_3\,(g) + 5\,O_2\,(g) \longrightarrow 4NO\,(g) + 6\,H_2O\,(l)$$

If we start with **10.0 g** of ammonia how much water can we produce?

1 mole NH_3 = 17.03 g

Converting mass of NH_3 to moles of NH_3:

$$10.0\,\cancel{g}\,NH_3 \times \frac{1\,mol\,NH_3}{17.03\,\cancel{g}} = 0.587\,mol\,NH_3$$

If we start with **10.0 g** of ammonia what mass of water can we produce?

$$4NH_3\,(g) + 5\,O_2\,(g) \longrightarrow 4NO\,(g) + 6\,H_2O\,(l)$$

4 mole NH_3 → 6 moles of H_2O

1 mole H_2O = 18.02 g

Converting moles NH_3 to grams H_2O:

$$0.587\,\cancel{\text{mol NH}_3} \times \frac{6\,\cancel{\text{mol H}_2\text{O}}}{4\,\cancel{\text{mol NH}_3}} \times \frac{18.02\text{ g H}_2\text{O}}{1\,\cancel{\text{mol H}_2\text{O}}} = \mathbf{15.9}\text{ g H}_2\text{O}$$

Determining the Limiting Reagent

$$4NH_3\,(g) \quad + \quad 5\,O_2\,(g) \longrightarrow 4NO\,(g) \quad + \quad 6\,H_2O\,(l)$$

If **10.0 *g*** of ammonia **and 10.0 *g* O_2** are used, what mass of water can we produce ?

1. Determine number of moles of each reactant using masses given:

$$10.0\,\cancel{g\,NH_3} \times \frac{1\,mol\,NH_3}{17.0\,\cancel{g\,NH_3}} = 0.587\,mol\,NH_3$$

$$10.0\,\cancel{g\,O_2} \times \frac{1\,mol\,O_2}{32.0\,\cancel{g\,O_2}} = 0.313\,mol\,O_2$$

$$4NH_3\,(g) + 5\,O_2\,(g) \longrightarrow 4NO\,(g) + 6\,H_2O\,(l)$$

If **10.0 g** of ammonia **and 10.0 g O_2** are used, what mass of water can we produce ?

2. Use balanced chemical equation to state number of moles that will **actually** be reacting in the chemical reaction.

4 moles of NH_3 reacts with 5 moles of O_2

$$0.587 \text{ mol } NH_3 \text{ reacts with } 0.587 \times \frac{5 \text{ mol } O_2}{4 \text{ mol } NH_3} = 0.734 \text{ mol } O_2$$

Moles in 10.0 g of NH_3

Moles of O_2 required to react with 10.0 g of NH_3

BUT the 10.0 g of O_2 given only contains 0.313 moles !

Oxygen will be completely consumed in the reaction !

This makes O_2 the limiting reagent

Dictates amount of products formed

1 mole O_2 = 32.0 g

5 moles $O_2 \rightarrow$ 6 moles H_2O

1 mole H_2O = 18.02 g

Converting moles of O_2 to mass of H_2O:

$$0.313 \cancel{\text{ mol } O_2} \times \frac{6 \cancel{\text{ mol } H_2O}}{5 \cancel{\text{ mol } O_2}} \times \frac{18.02 \text{ g}}{1 \cancel{\text{ mol } H_2O}} = \mathbf{6.77 \text{ g } H_2O}$$

theoretical yield

In performing this experiment only 4.00 g of water was produced.

What is the percentage yield of water produced ?

Experimental/ actual yield: 4.00 g

Theoretical yield : 6.77 g

$$\% \text{ yield} = \frac{\text{actual}}{\text{theoretical}} = \frac{4.00}{6.77} \times 100 = 59.1\,\%$$