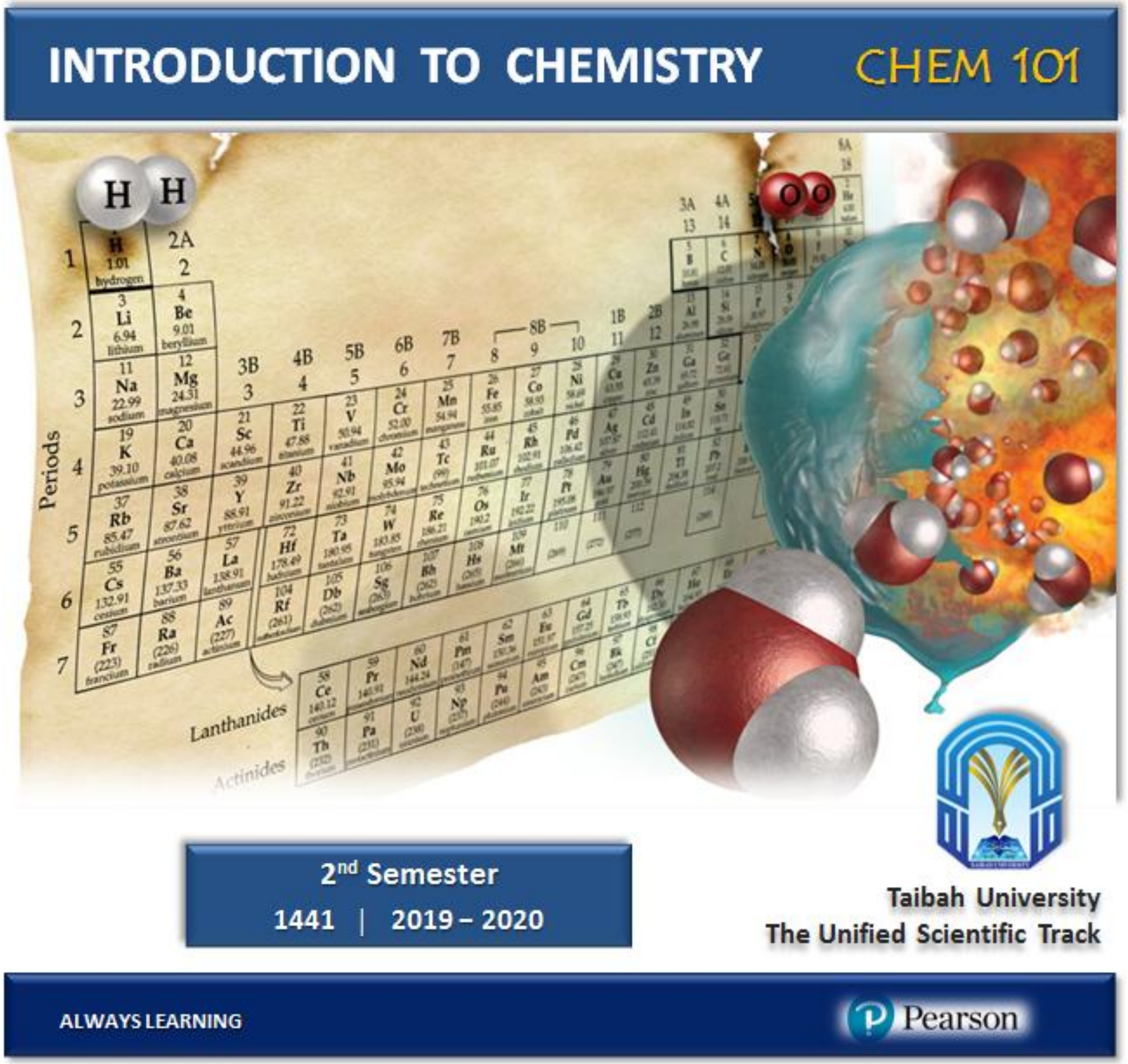

INTRODUCTION TO CHEMISTRY CHEM 101

2nd Semester
1441 | 2019 – 2020

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Lecture Presentation

Chapter 4

Stoichiometry, Solution Concentration and Chemical Reactions

Topic 13

- Reaction Stoichiometry
- Limiting Reactant, Theoretical Yield & Percent Yield

What Is Meant By Stoichiometry?

- **Stoichiometry**: calculations of the quantities of reactants and products in a chemical reaction.
  - ✓ Stoichiometry allows us to predict the amounts of products that will form in a chemical reaction based on the amount of reactants.
  - ✓ Stoichiometry also allows us to determine the amount of reactants necessary to form a given amount of product.

4.1 Reaction Stoichiometry: How Much CO_2 is Produced?

- The coefficients in a balanced chemical equation specify the relative amounts in moles of each of the substances involved in the reaction:

Example: $2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$

2 molecules of C_8H_{18} react with 25 molecules of O_2 to form 16 molecules of CO_2 and 18 molecules of H_2O.

Or: 2 moles of C_8H_{18} react with 25 moles of O_2 to form 16 moles of CO_2 and 18 moles of H_2O.

2 mol C_8H_{18} : 25 mol O_2 : 16 mol CO_2 : 18 mol H_2O

Making Molecules: Mole-to-Mole Conversions

From the balanced equation of the combustion of octane:

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

we can write the following stoichiometric ratio:

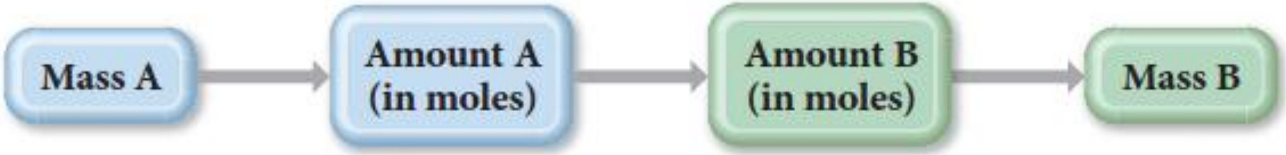
2 moles $C_8H_{18}(l)$: 16 moles CO_2

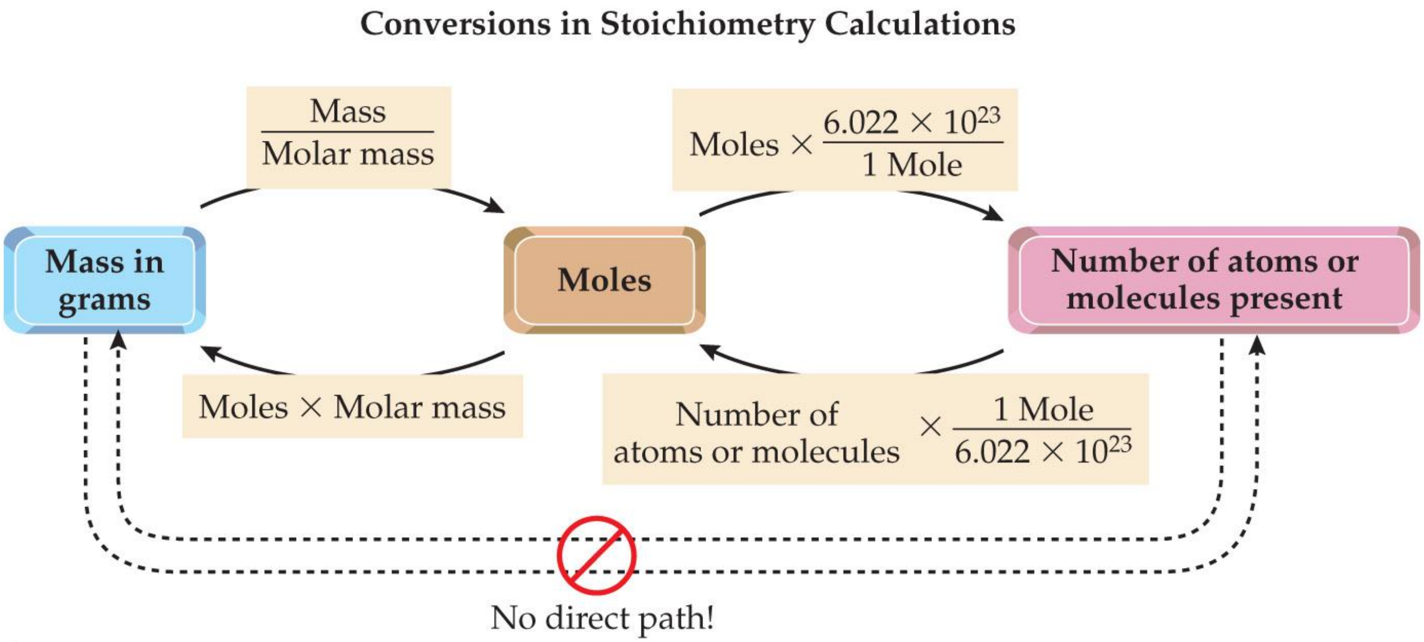
(This ratio is called: The Conversion Factor)

Suppose that we burn 22 moles of **C_8H_{18}**: the amount of **CO_2** produced can be calculated using the conversion factor, as follows:

$$\frac{22\ moles\ of\ C_8H_{18} \times 16\ moles\ of\ CO_2}{2\ moles\ of\ C_8H_{18}} = 176\ moles\ of\ CO_2$$

Making Molecules: Mass-to-Mass Conversions




Conversions in Stoichiometry Calculations

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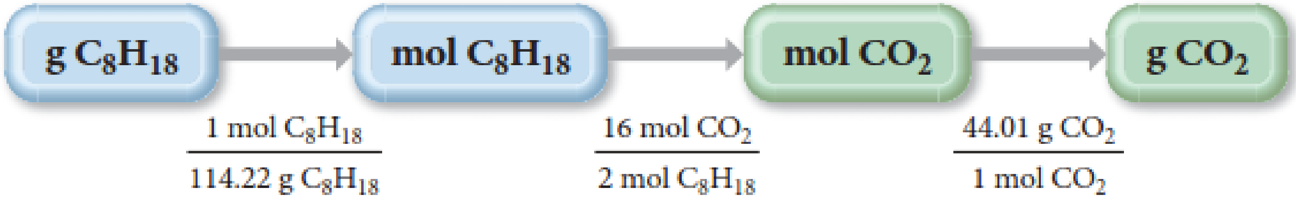
Example 1: Estimate the mass of CO_2 emitted into the atmosphere in 2010 by the combustion of 3.5×10^{15} g gasoline.

$$2\ C_8H_{18}(l) + 25\ O_2(g) \longrightarrow 16\ CO_2(g) + 18\ H_2O(g)$$

Given: **Find:**	3.4×10^{15} g C_8H_{18} g CO_2
Plan:	
Relationships:	1 mol C_8H_{18} = 114.22g, 1 mol CO_2 = 44.01g, 2 mol C_8H_{18}:16 mol CO_2

Solution:

$$3.5 \times 10^{15}\ \cancel{\text{g } C_8H_{18}} \times \frac{1\ \cancel{\text{mol } C_8H_{18}}}{114.22\ \cancel{\text{g } C_8H_{18}}} \times \frac{16\ \cancel{\text{mol } CO_2}}{2\ \cancel{\text{mol } C_8H_{18}}} \times \frac{44.01\text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}} = 1.1 \times 10^{16}\text{ g } CO_2$$

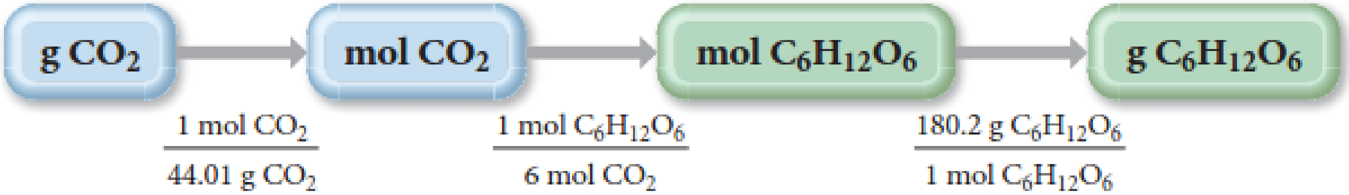
Example 2: How many grams of glucose can be synthesized from 37.8 g of CO_2 in photosynthesis?

$$6\,CO_2(g) + 6\,H_2O(l) \xrightarrow{\text{sunlight}} 6\,O_2(g) + C_6H_{12}O_6(aq)$$

Given: 37.8 g CO_2

Find: g $C_6H_{12}O_6$

Plan:



$$\text{g } CO_2 \xrightarrow[\frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2}]{} \text{mol } CO_2 \xrightarrow[\frac{1 \text{ mol } C_6H_{12}O_6}{6 \text{ mol } CO_2}]{} \text{mol } C_6H_{12}O_6 \xrightarrow[\frac{180.2 \text{ g } C_6H_{12}O_6}{1 \text{ mol } C_6H_{12}O_6}]{} \text{g } C_6H_{12}O_6$$

Relationships: 1 mol $C_6H_{12}O_6$ = 180.2 g, 1 mol CO_2 = 44.01 g, 1 mol $C_6H_{12}O_6$: 6 mol CO_2

Solution:

$$37.8\ \cancel{\text{g } CO_2} \times \frac{1\ \cancel{\text{mol } CO_2}}{44.01\ \cancel{\text{g } CO_2}} \times \frac{1\ \cancel{\text{mol } C_6H_{12}O_6}}{6\ \cancel{\text{mol } CO_2}} \times \frac{180.16 \text{ g } C_6H_{12}O_6}{1\ \cancel{\text{mol } C_6H_{12}O_6}} = 25.8 \text{ g } C_6H_{12}O_6$$

Assessment

1. Calculate the number of NO_2 moles that will be formed when each amount of N_2O_5 completely dissociates:

$$2\ N_2O_5(g) \rightarrow 4\ NO_2(g) + O_2(g)$$

a) 1.3 mol of N_2O_5 **b)** 1.55 kg of N_2O_5 **c)** 10.5 g of N_2O_5 **d)** 2.25×10^{23} molecules of N_2O_5

2. How many moles of H_2O would be produced when 5 moles of C_2H_6O completely react with oxygen gas according to the equation?

$$C_2H_6O + 3\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O$$

3. What is the mass (in g) of $AICl_3$ that will be produced when 95 grams of Al completely react with excess Cl_2 according to this equation?

$$2\ Al + 3\ Cl_2 \rightarrow 2\ AICl_3$$

4. How many moles of CO_2 would be produced when 4.5×10^{23} molecules of C_3H_7COOH completely react with oxygen gas according to the following equation?

$$C_3H_7COOH + 5\ O_2 \rightarrow 4\ CO_2 + 4\ H_2O$$

5. Lithium and nitrogen react to produce lithium nitride as follows: $6\ Li(s) + N_2(g) \rightarrow 2\ Li_3N(s)$

How many grams of N_2 are needed to fully react with 15 g of lithium?

6. Given the following reaction: $N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$

- **a)** How many grams of N_2 are required by 35 g of H_2 to make a complete reaction?
- **b)** What is the mass (in g) of NH_3 that will be produced from 35 g of H_2?

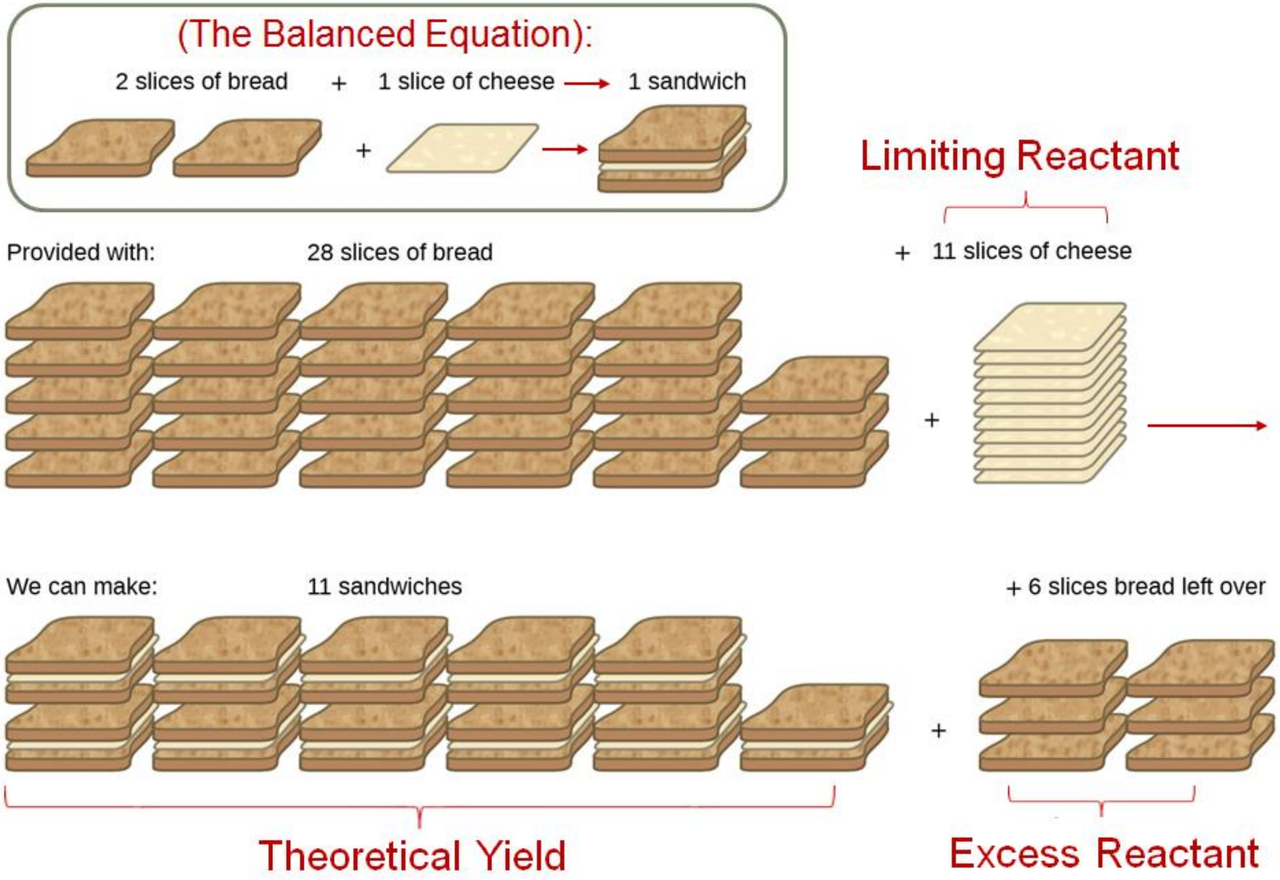
4.2 Limiting Reactant and Theoretical Yield

Limiting Reactant: Consider this food analogy, making cheese sandwiches:

2 slices of bread + 1 slice of cheese = 1 sandwich ….. (the equation)

Starting with 28 slices of bread and 11 slices of cheese, one may prepare 11 sandwiches.

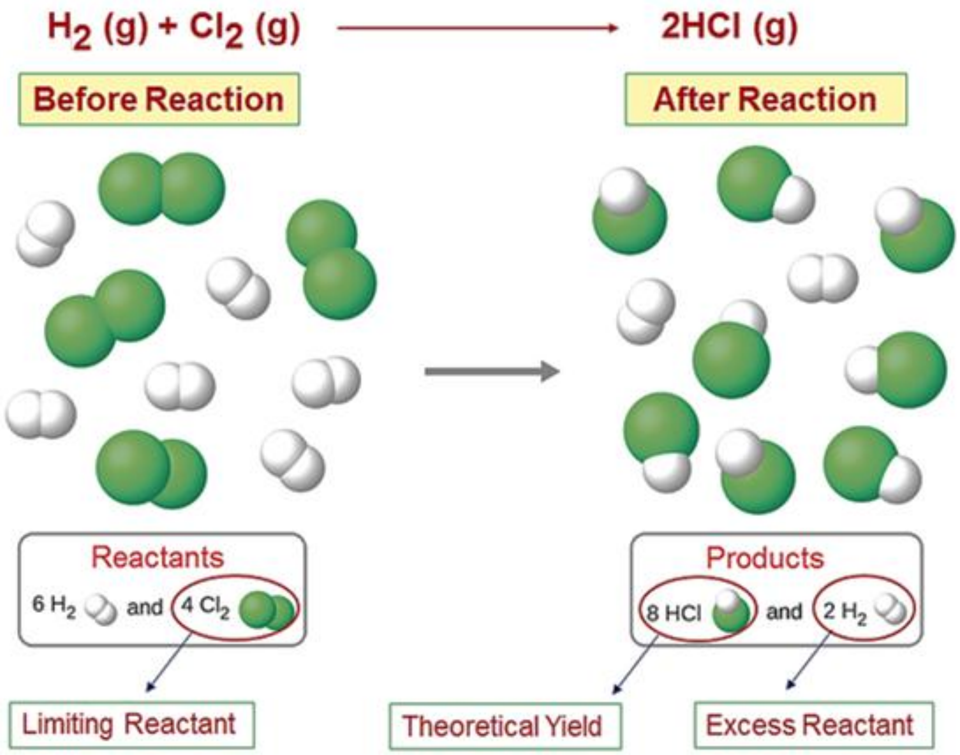
In this scenario, the number of sandwiches prepared has been **limited** by the number of cheese slices, and the bread slices have been provided in **excess**.



4.2 Limiting Reactant and Theoretical Yield

Now, consider the same concept with regard to a chemical reaction:

For the following reaction, if we started with **$6H_2$**, and **$4Cl_2$**, find: Limiting Reactant, Excess Reactant, and Theoretical Yield?



Limiting Reactant and Theoretical Yield: Summary

- **Limiting Reactant**: is the reactant that is completely consumed in a chemical reaction and limits the amount of product.
  - ✓ Reactant that is in short supply
- **Excess Reactant**: is any reactant that occurs in a quantity greater than is required to completely react with the limiting reactant.
  - ✓ Present in excess, will be left over after the reaction
- **Theoretical Yield**: is the calculated amount of product that can be made in a chemical reaction based on the amount of the limiting reactant.
- **Actual Yield**: is the amount of product actually produced in a chemical reaction.
  - ✓ Less than theoretical yield
- **Percent Yield**:

$$\boxed{\textbf{The percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%}$$

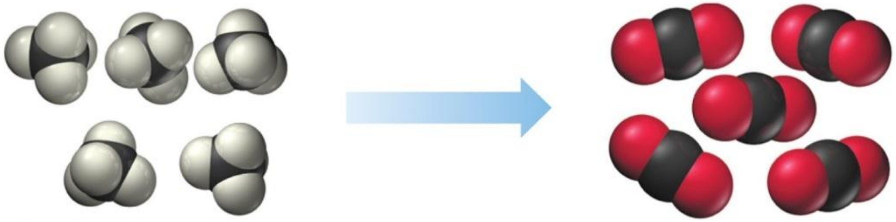
Limiting Reactant and Theoretical Yield: Exercise

- If we have five molecules of CH_4 and eight molecules of O_2, which is the limiting reactant?

$$CH_4(g) + 2\,O_2(g) \longrightarrow CO_2(g) + 2\,H_2O(l)$$

✓ First, we calculate the number of CO_2 molecules that can be made from 5 CH_4 molecules.

$$5\,\cancel{CH_4} \times \frac{1\,CO_2}{1\,\cancel{CH_4}} = 5\,CO_2$$



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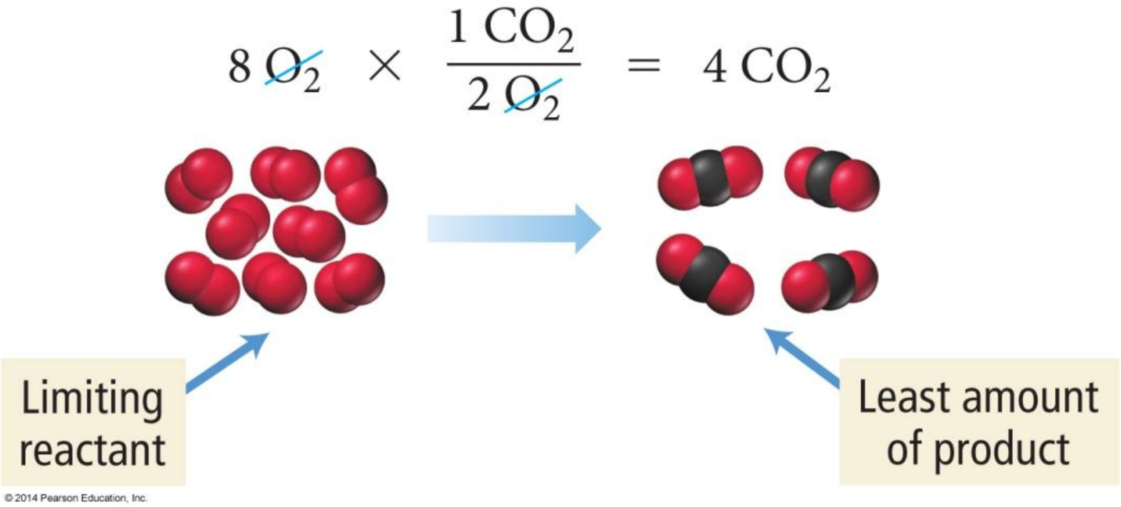
Limiting Reactant and Theoretical Yield: Exercise

- Then, we calculate the number of CO_2 molecules that can be made from 8 O_2 molecules:



$$8\ \cancel{O_2} \times \frac{1\ CO_2}{2\ \cancel{O_2}} = 4\ CO_2$$

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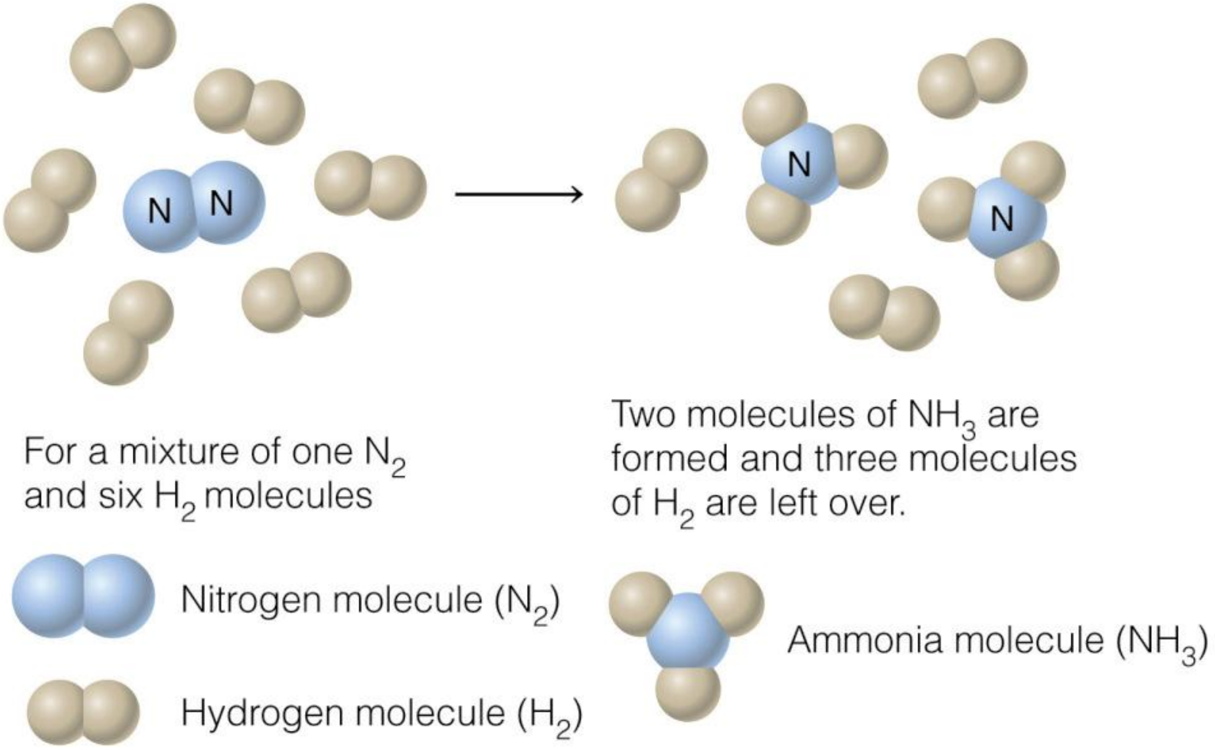
- We have enough CH_4 to make 5 CO_2 molecules and enough O_2 to make 4 CO_2 molecules.
- Therefore, O_2 is the limiting reactant, and
- 4 CO_2 molecules is the theoretical yield (based on limiting reactant).
- CH_4 is the reactant in excess (or, the excess reactant).

Limiting Reactant and Theoretical Yield: Exercise

- In the following reaction, determine the limiting reactant, the excess reactant, and the theoretical yield if **1 mole N_2** reacted with **6 moles H_2**, according to the following balanced equation:

$$N_2(g) + 3\ H_2(g) \rightarrow 2\ NH_3(g)$$



For a mixture of one N_2 and six H_2 molecules

Two molecules of NH_3 are formed and three molecules of H_2 are left over.

Nitrogen molecule (N_2)

Hydrogen molecule (H_2)

Ammonia molecule (NH_3)

Calculating The Percent Yield

- What is the **percent yield** if **225.10 g** of **HI** were isolated out of a possible yield of **255.824 g HI**?

Given: actual yield = 225.10 g, and

theoretical yield = 255.824 g

$$\boxed{\textbf{The percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%}$$

$$\%\,\text{yield} = \frac{225.10\ \text{g HI}}{255.826\ \text{g HI}} \times 100 = 87.989\%$$

Hint: if both Actual and theoretical yields are given in the same unit (grams or moles), then no need to convert any of them. **But**, if one of them is in grams and the other is in moles, then you must convert one of them first!

Limiting Reactant and Percent Yield: Gram-to-Gram

Example: $2\ \text{Na}(s) + \text{Cl}_2(g) \rightarrow 2\ \text{NaCl}(s)$

If we have: 53.2 **g** of Na, and 65.8 **g** of Cl_2

Find: a) The limiting reactant and theoretical yield (in g)

b) If the actual yield = 86.4 g NaCl, calculate % yield.

Solution:

$$\left.\begin{aligned}
&53.2\ \cancel{\text{g Na}} \times \frac{1\ \cancel{\text{mol Na}}}{22.99\ \cancel{\text{g Na}}} \times \frac{2\ \cancel{\text{mol NaCl}}}{2\ \cancel{\text{mol Na}}} \times \frac{58.44\ \text{g NaCl}}{1\ \cancel{\text{mol NaCl}}} = 135\ \text{g NaCl} \\
&65.8\ \cancel{\text{g Cl}_2} \times \frac{1\ \cancel{\text{mol Cl}_2}}{70.90\ \cancel{\text{g Cl}_2}} \times \frac{2\ \cancel{\text{mol NaCl}}}{1\ \cancel{\text{mol Cl}_2}} \times \frac{58.44\ \text{g NaCl}}{1\ \cancel{\text{mol NaCl}}} = 108\ \text{g NaCl}
\end{aligned}\right\} \text{Smallest amount determines limiting reactant.}$$

Limiting reactant

Least amount of product

$$\boxed{\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% = \frac{86.4\ \text{g}}{108\ \text{g}} \times 100\% = 80.0\%}$$

Limiting Reactant and Percent Yield: Gram-to-Gram

EXAMPLE 4.2 Limiting Reactant and Theoretical Yield

Ammonia, NH_3, can be synthesized by the reaction:

$$2\,NO(g) + 5\,H_2(g) \longrightarrow 2\,NH_3(g) + 2\,H_2O(g)$$

Starting with 86.3 g NO and 25.6 g H_2, find the theoretical yield of ammonia in grams.

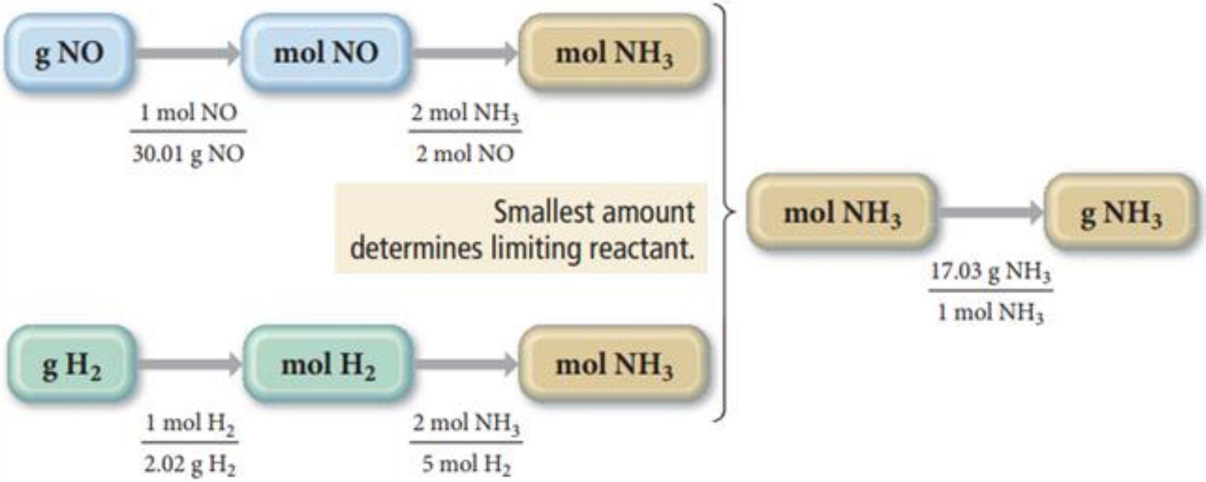
SORT You are given the mass of each reactant in grams and asked to find the theoretical yield of a product.

GIVEN 86.3 g NO, 25.6 g H_2

FIND theoretical yield of NH_3

STRATEGIZE Determine which reactant makes the least amount of product by converting from grams of each reactant to moles of the reactant to moles of the product. Use molar masses to convert between grams and moles and use the stoichiometric relationships (deduced from the chemical equation) to convert

CONCEPTUAL PLAN



$$\text{g NO} \xrightarrow{\frac{1\text{ mol NO}}{30.01\text{ g NO}}} \text{mol NO} \xrightarrow{\frac{2\text{ mol }NH_3}{2\text{ mol NO}}} \text{mol }NH_3$$

$$\text{g }H_2 \xrightarrow{\frac{1\text{ mol }H_2}{2.02\text{ g }H_2}} \text{mol }H_2 \xrightarrow{\frac{2\text{ mol }NH_3}{5\text{ mol }H_2}} \text{mol }NH_3$$

Smallest amount determines limiting reactant.

$$\text{mol }NH_3 \xrightarrow{\frac{17.03\text{ g }NH_3}{1\text{ mol }NH_3}} \text{g }NH_3$$

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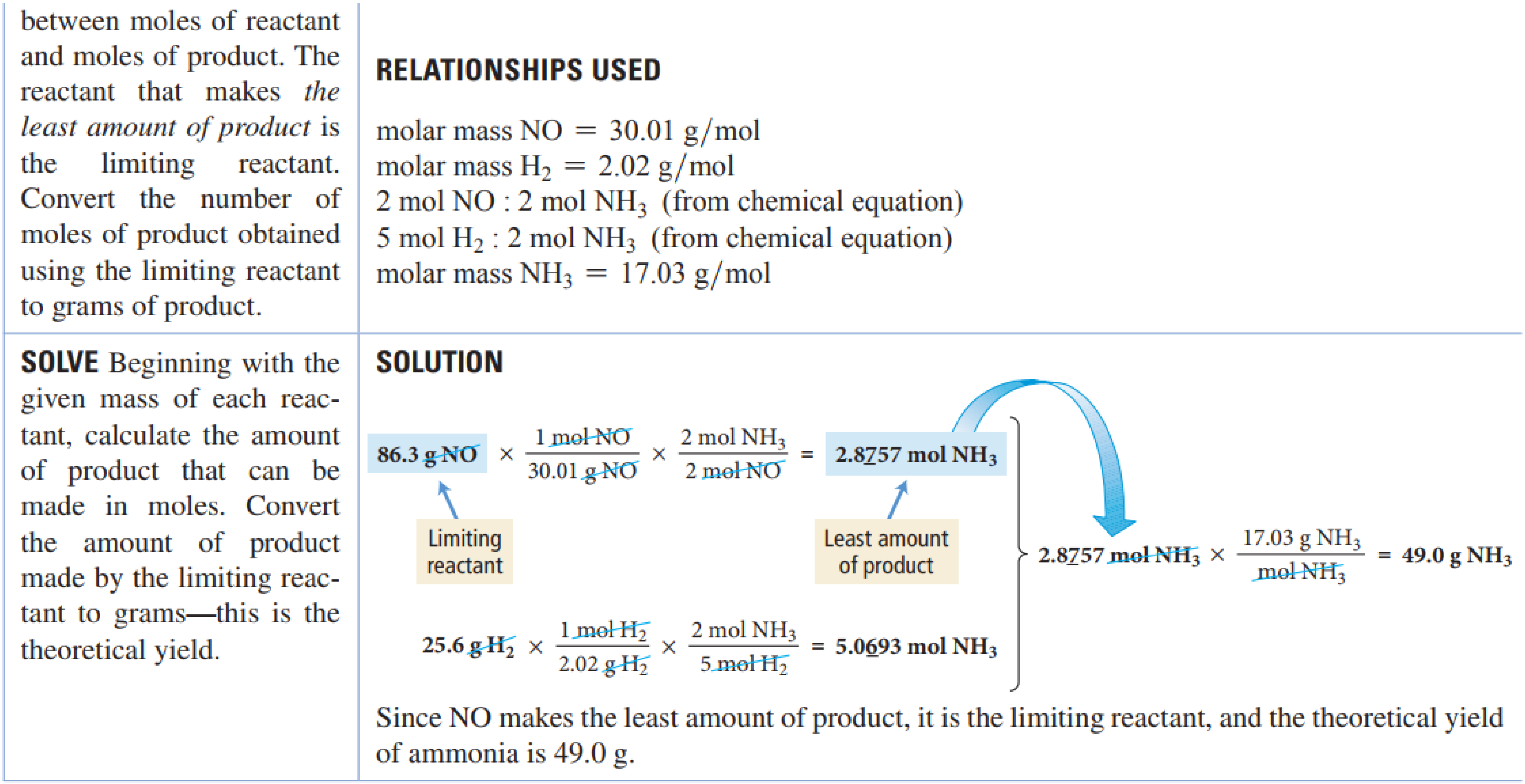


between moles of reactant and moles of product. The reactant that makes *the least amount of product* is the limiting reactant. Convert the number of moles of product obtained using the limiting reactant to grams of product.

RELATIONSHIPS USED

molar mass NO = 30.01 g/mol
molar mass H_2 = 2.02 g/mol
2 mol NO : 2 mol NH_3 (from chemical equation)
5 mol H_2 : 2 mol NH_3 (from chemical equation)
molar mass NH_3 = 17.03 g/mol

SOLVE Beginning with the given mass of each reactant, calculate the amount of product that can be made in moles. Convert the amount of product made by the limiting reactant to grams—this is the theoretical yield.

SOLUTION

$$\mathbf{86.3\ \cancel{g\ NO}} \times \frac{1\ \cancel{\text{mol NO}}}{30.01\ \cancel{\text{g NO}}} \times \frac{2\ \text{mol NH}_3}{2\ \cancel{\text{mol NO}}} = \mathbf{2.8\underline{7}57\ mol\ NH_3}$$

$$\mathbf{25.6\ \cancel{g\ H_2}} \times \frac{1\ \cancel{\text{mol H}_2}}{2.02\ \cancel{\text{g H}_2}} \times \frac{2\ \text{mol NH}_3}{5\ \cancel{\text{mol H}_2}} = \mathbf{5.0\underline{6}93\ mol\ NH_3}$$

$$\mathbf{2.8\underline{7}57\ \cancel{mol\ NH_3}} \times \frac{17.03\ \text{g NH}_3}{\cancel{\text{mol NH}_3}} = \mathbf{49.0\ g\ NH_3}$$

Since NO makes the least amount of product, it is the limiting reactant, and the theoretical yield of ammonia is 49.0 g.

Keep in mind: before working on stoichiometric calculations, make sure that:

1- The chemical equation is **balanced**.

2- The amounts of substances are in “**moles**”. If they are given in “grams” or Kg, convert them first to “moles”.

Assessment

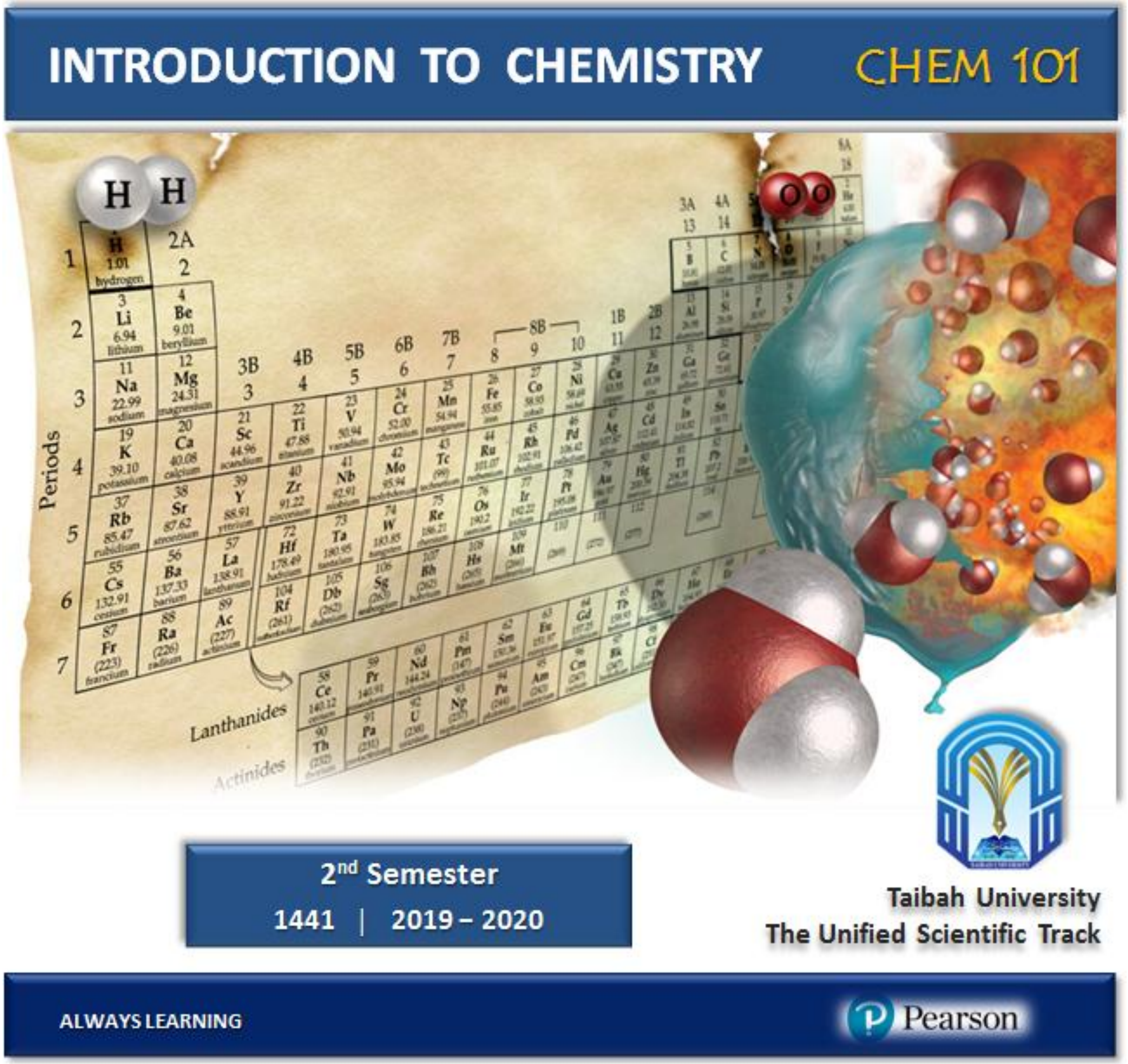
1- For the following reaction, find the **limiting reactant, excess reactant, and theoretical yield (in moles)** if we started the reaction with 12.6 **mol** Na and 6.9 **mol** Br_2

$$2\ Na(s) + Br_2(g) \rightarrow 2\ NaBr(s)$$

2- For the following reaction, calculate the **theoretical yield** of product (in g) if we started the reaction with 7.5 **g** Al and 24.8 **g** Cl_2

$$2\ Al(s) + 3\ Cl_2(g) \rightarrow 2\ AlCl_3(s)$$

3- What is the **percent yield** for a reaction if its theoretical yield is 83 **g** and its actual yield is 75 **g**?



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Chapter 4

Stoichiometry, Solution Concentration and Chemical Reactions

Topic 14

- Solution Concentration
- Types of Aqueous Solutions

4.3 Concentration of Solutions

What Is a "Solution"?

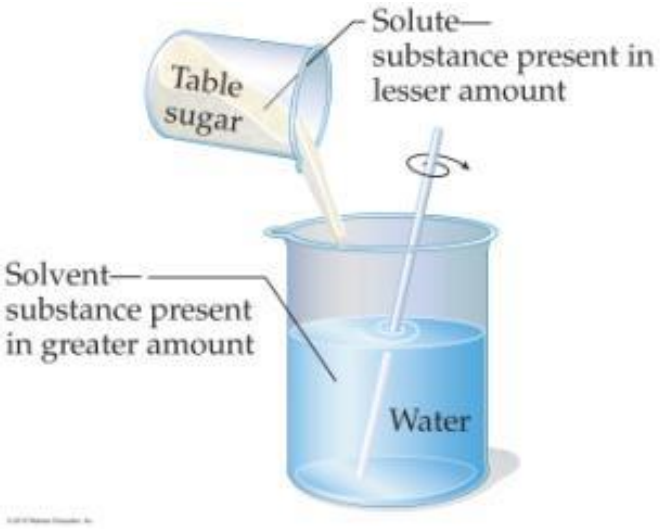
Solution: A homogenous mixture of two or more substances:

- **Solvent**: material present in largest amount.
- **Solute**: all other materials present.

- Example:

Consider sugar dissolved in water:

- Water is the solvent.
- Sugar is the solute.



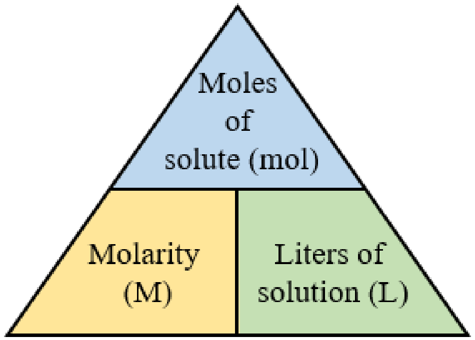
4.3 Concentration of Solutions

- **Concentration:** is the amount of solute present in the solution.

- **Molarity:** is a method to express the concentration. It shows the relationship between the moles of solute and liters of solution.

$$\text{Molarity (M)} = \frac{\text{amount of solute (in mol)}}{\text{volume of solution (in L)}}$$

$$M = \frac{n}{V}$$



✓ Unit of molarity **(M)** = moles of solute / liter of solution

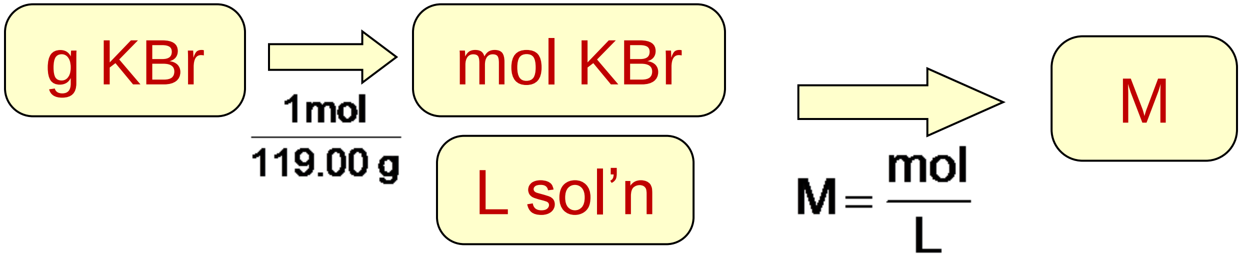
$$\mathbf{M = mol/L = mol.L^{-1} = molar}$$

Example: Find the molarity of a solution that has 25.5 g KBr dissolved in 1.75 L of solution

Given: 25.5 g KBr, 1.75 L solution

Find: molarity, M

Plan:



$$\begin{tikzcd} \text{g KBr} \arrow[r, "\frac{1\,\text{mol}}{119.00\,\text{g}}"'] & \text{mol KBr},\ \text{L sol'n} \arrow[r, "M = \frac{\text{mol}}{\text{L}}"'] & \text{M} \end{tikzcd}$$

Relationships: 1 mol KBr = 119.00 g, M = moles/L

Solution:

$$25.5\ \cancel{\text{g KBr}} \times \frac{1\ \text{mol KBr}}{119.00\ \cancel{\text{g KBr}}} = 0.21\underline{4}29\ \text{mol KBr}$$

$$\text{molarity, M} = \frac{\text{moles KBr}}{\text{L solution}} = \frac{0.21\underline{4}29\ \text{mol KBr}}{1.75\ \text{L}} = 0.122\ \text{M}$$

Check: because most solutions are between 0 and 18 M, the answer makes sense

Example: How many liters of 0.125 M NaOH contain 0.255 mol NaOH?

Given: 0.125 M NaOH, 0.255 mol NaOH

Find: liters, L

Plan:

$$\text{mol NaOH} \xrightarrow[\frac{\text{1 L solution}}{\text{0.125 mol NaOH}}]{} \text{L sol'n}$$

Relationships: 0.125 mol NaOH = 1 L solution

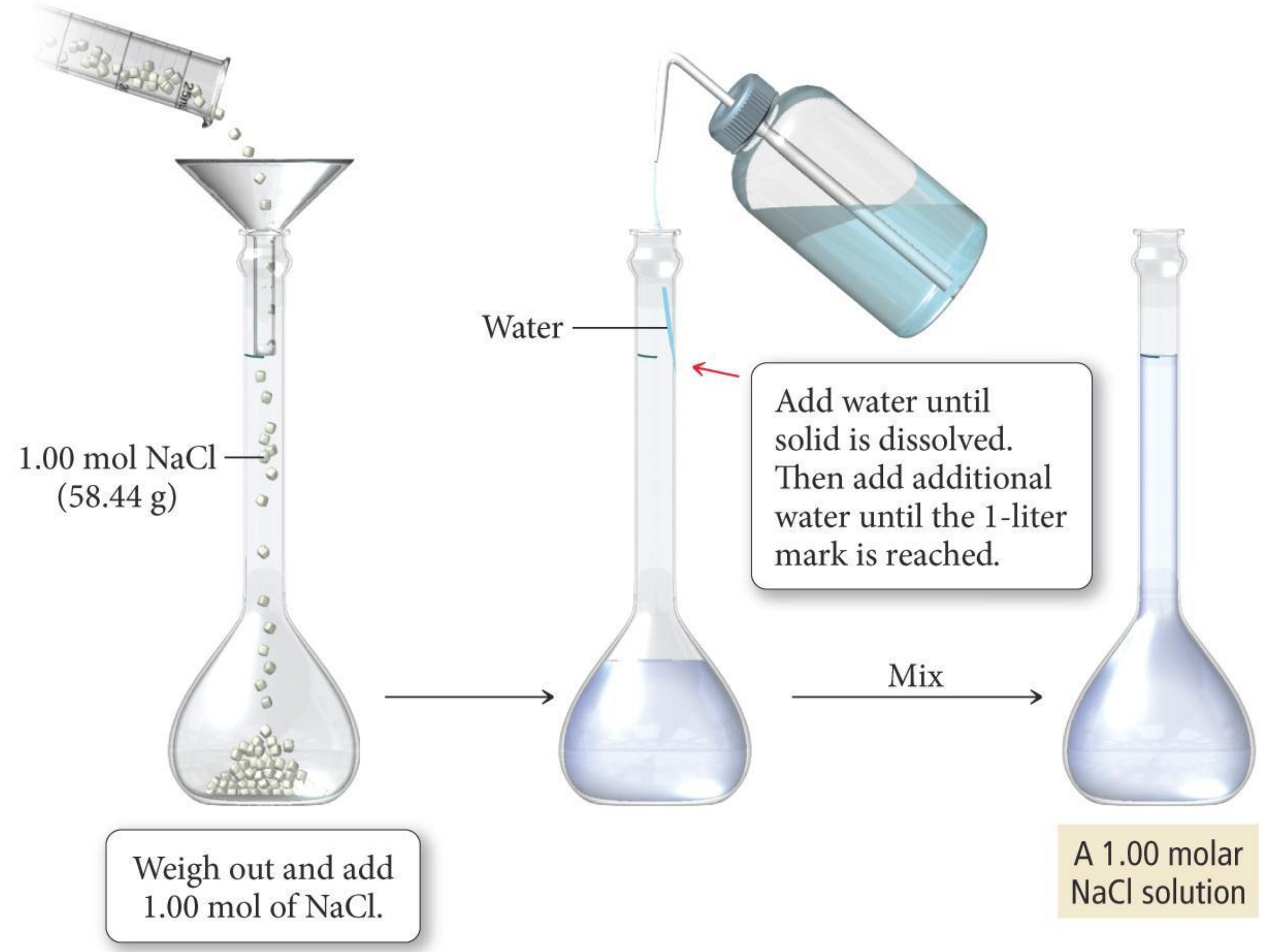
Solution:

$$0.255\ \text{mol}\ \cancel{\text{NaOH}} \times \frac{\text{1 L solution}}{0.125\ \text{mol}\ \cancel{\text{NaOH}}} = 2.04\ \text{L solution}$$

Check: because each L has only 0.125 mol NaOH, it makes sense that 0.255 mol should require a little more than 2 L

Example: Preparing 1 L of a 1 M NaCl Solution

Preparing a Solution of Specified Concentration



Weigh out and add 1.00 mol of NaCl.

Add water until solid is dissolved. Then add additional water until the 1-liter mark is reached.

A 1.00 molar NaCl solution

Solution Dilution

- Often, solutions are stored as concentrated stock solutions.
- To make solutions of lower concentrations from these stock solutions, more solvent is added.
  - The amount of solute doesn’t change, just the volume of solution:

solute moles in solution 1 = solute moles in solution 2

- The concentrations and volumes of the stock and new solutions are inversely proportional:

$$\boxed{M_1V_1 = M_2V_2}$$

Solution Dilution: Example

Suppose a laboratory procedure requires **3.00 L** of a **0.500 M** CaCl2 solution. How should we prepare this solution from a **10.0 M** stock solution?

Answer:

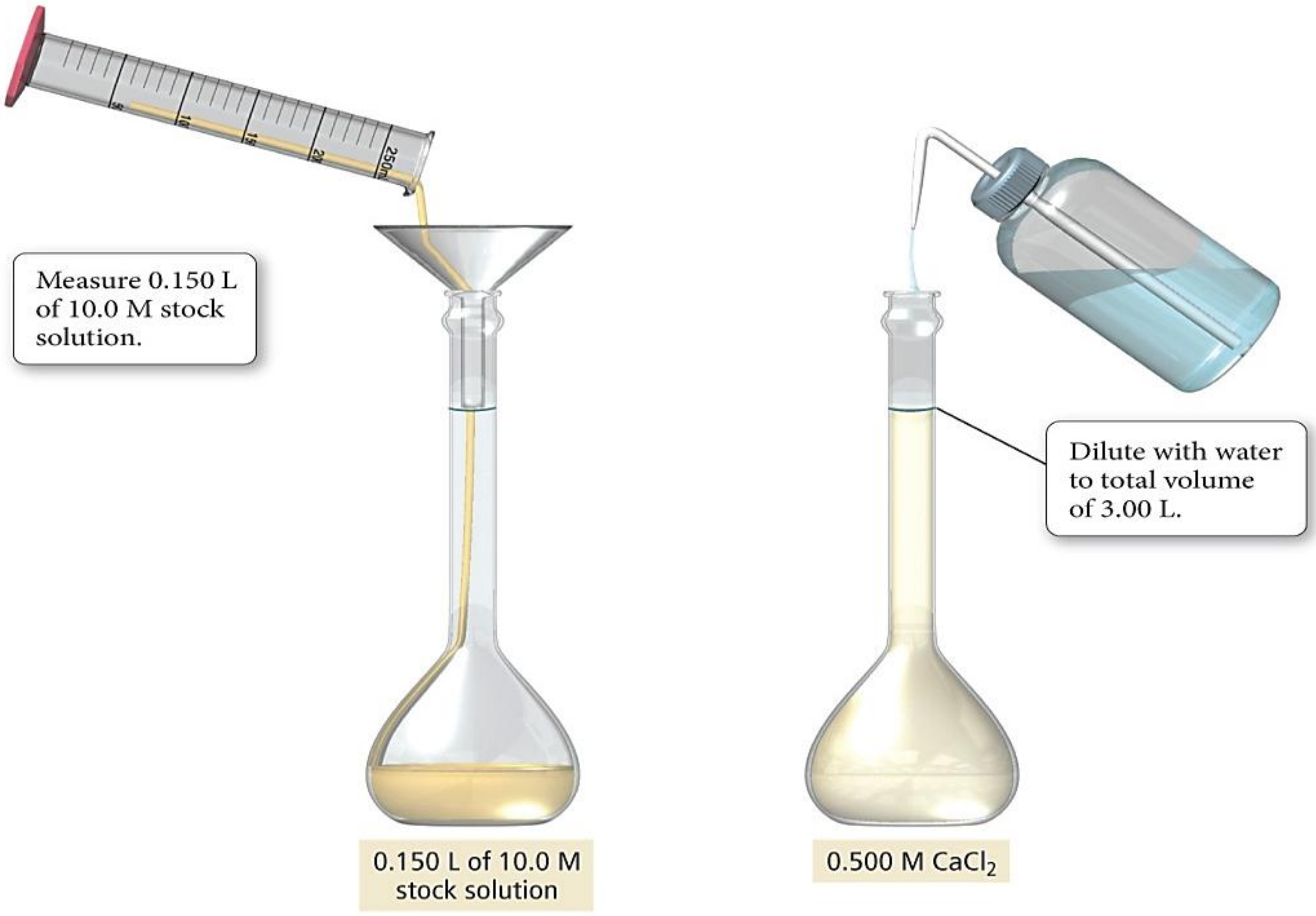
$$V_1 = \frac{M_2V_2}{M_1}$$

$$= \frac{0.500\ \cancel{\text{mol/L}} \times 3.00\ \text{L}}{10.0\ \cancel{\text{mol/L}}}$$

$$= 0.150\ \text{L}$$

Example: How to Prepare 3.00 L of 0.500 M $CaCl_2$ from a 10.0 M Stock Solution?

Diluting a Solution



Measure 0.150 L of 10.0 M stock solution.

Dilute with water to total volume of 3.00 L.

0.150 L of 10.0 M stock solution

0.500 M $CaCl_2$

$$M_1V_1 = M_2V_2$$

$$\frac{10.0\ \text{mol}}{\cancel{\text{L}}} \times 0.150\cancel{\text{L}} = \frac{0.500\ \text{mol}}{\cancel{\text{L}}} \times 3.00\cancel{\text{L}}$$

$$1.50\ \text{mol} = 1.50\ \text{mol}$$

Practice: How would you prepare 200.0 mL of 0.25 M NaCl solution from a 2.0 M solution?

Given: $M_1 = 2.0\ \text{M}$, $M_2 = 0.25\ \text{M}$, $V_2 = 200.0\ \text{mL}$

Find: V_1 (in mL)

Plan:

$$\boxed{M_1, M_2, V_2} \xrightarrow{} \boxed{V_1}$$

$$V_1 = \frac{M_2 \bullet V_2}{M_1}$$

Relationships: $M_1V_1 = M_2V_2$

Solution:

$$\frac{\left(0.25\,\frac{\cancel{\text{mol}}}{\text{L}}\right)\bullet\left(200.0\ \text{mL}\right)}{\left(2.0\,\frac{\cancel{\text{mol}}}{\text{L}}\right)} = 25\ \text{mL}$$

Dilute 25 mL of 2.0 M solution up to 200.0 mL

Check: because the solution is diluted by a factor of 8, the volume should increase by a factor of 8, and it does

Assessment

1- Calculate the molarity of each solution:

a) 4.3 mol of LiCl in 2.75 L solution.

b) 21.5 g $C_6H_{12}O_6$ in 1.85 L of solution.

2- How many moles of KCl are there in each solution?

a) 0.55 L of a 2.3 M KCl solution.

b) 114 mL of a 1.85 M KCl solution.

3- A saline solution contains 1.5 g of sodium chloride, NaCl, dissolved in 100 mL of solution. What is the molar concentration of the solution?

4- A laboratory procedure calls for making 400 mL of a 1.3 M $NaNO_3$ solution. What mass of $NaNO_3$ (in g) is needed?

5- If 123 mL of a 1.1 M glucose solution is diluted to 500 mL, what is the molarity of the diluted solution?

6- To what volume should you dilute 50 mL of a 12 M stock HNO_3 solution to obtain a 0.1 M HNO_3 solution?

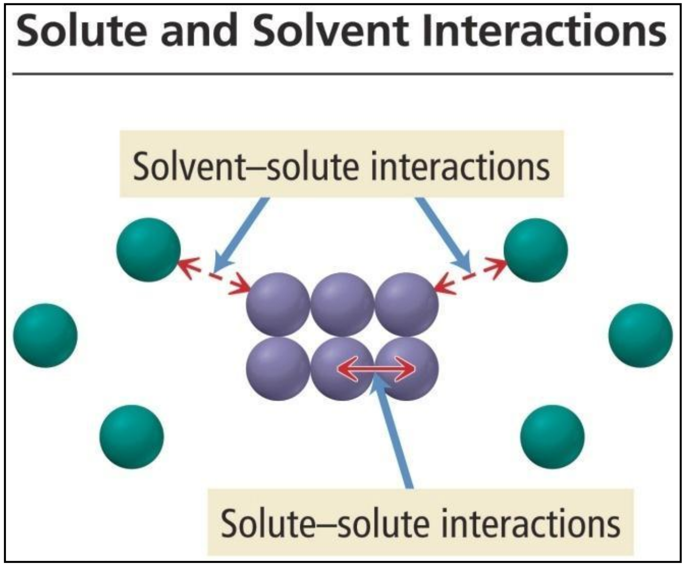
4.4 Types of Aqueous Solutions and Solubility

- Consider two familiar aqueous solutions: salt water and sugar water:
  - Salt water is a homogeneous mixture of NaCl and H_2O.
  - Sugar water is a homogeneous mixture of $C_{12}H_{22}O_{11}$ and H_2O.
- As you stir either of these two substances into the water, it seems to disappear.
  - How do solids such as salt and sugar dissolve in water?

What Happens When a Solute Dissolves?

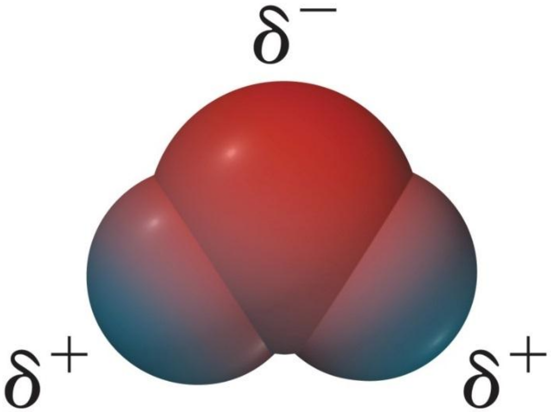
- There are attractive forces between the solute particles holding them together.
- There are also attractive forces between the solvent molecules.

- When we mix the solute with the solvent, there are attractive forces between the solute particles and the solvent molecules.
- If the attractions between solute and solvent are strong enough, the solute will dissolve.



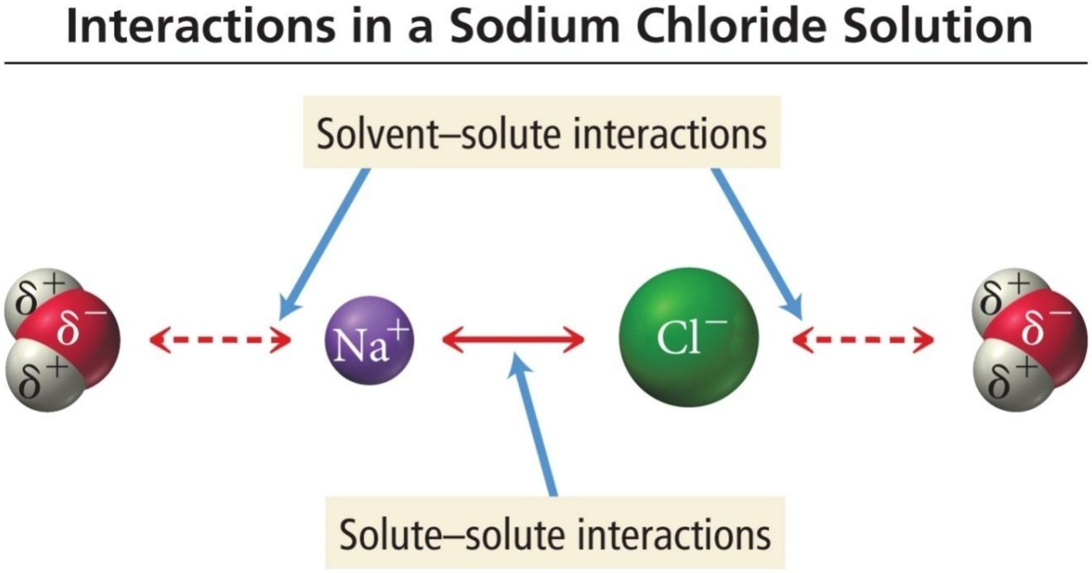
Charge Distribution in a Water Molecule

- There is an uneven distribution of electrons within the water molecule.
  - This causes the oxygen side of the molecule to have a partial negative charge (δ^-) and the hydrogen side to have a partial positive charge (δ^+).



Solute and Solvent Interactions in a Sodium Chloride Solution (or Other Ionic Compounds)

- When sodium chloride is put into water, the attraction of Na^+ and Cl^- ions to water molecules competes with the attraction among the oppositely charged ions themselves.



Interactions in a Sodium Chloride Solution

Solvent–solute interactions

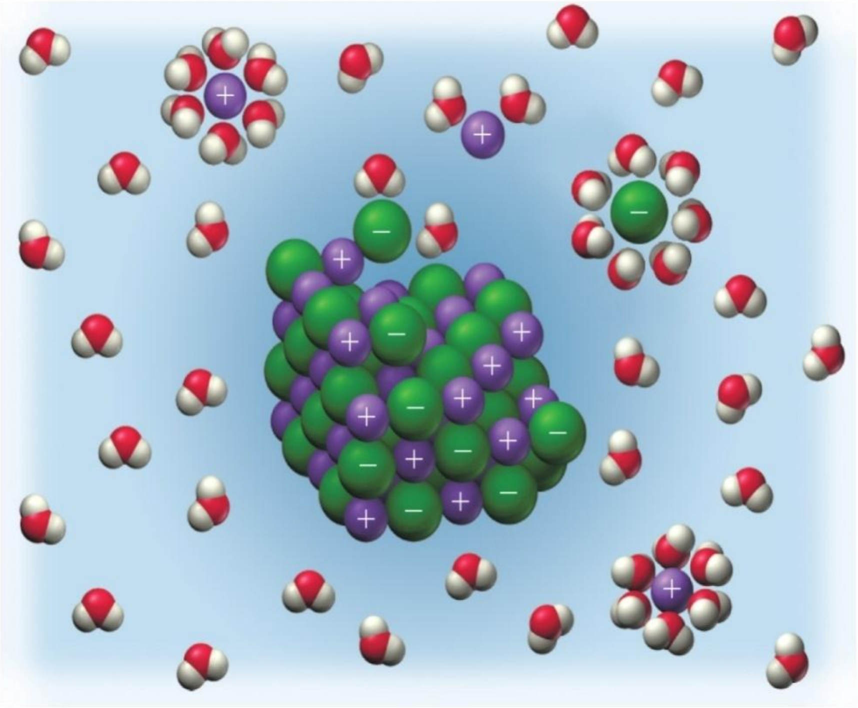
Solute–solute interactions

Dissolving Sodium Chloride in Water

- ✓ Each ion is attracted to the surrounding water molecules and pulled off and away from the crystal.

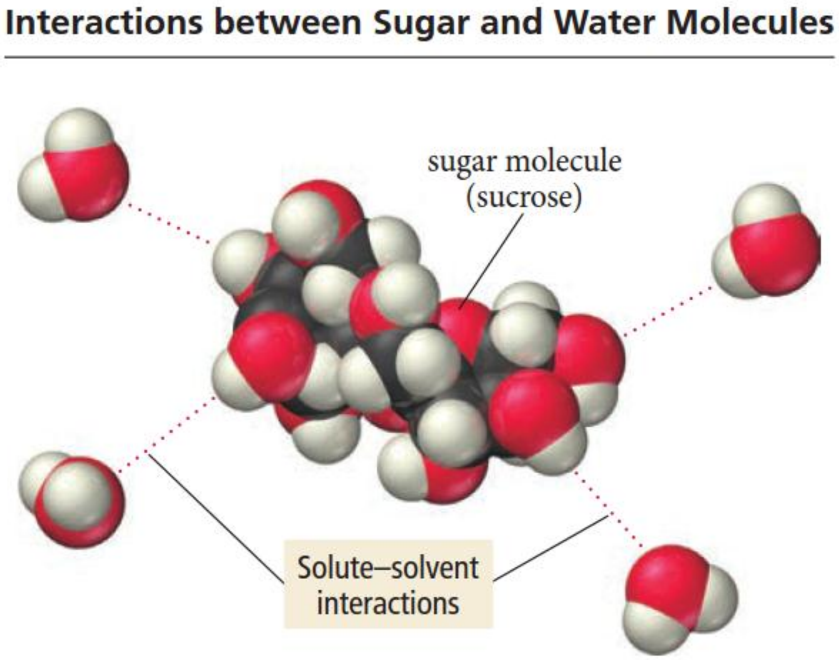
- ✓ Compounds such as salt that dissociate into ions when dissolved in water are called electrolytes, and the resulting solutions is able to conduct electricity.

Dissolution of an Ionic Compound



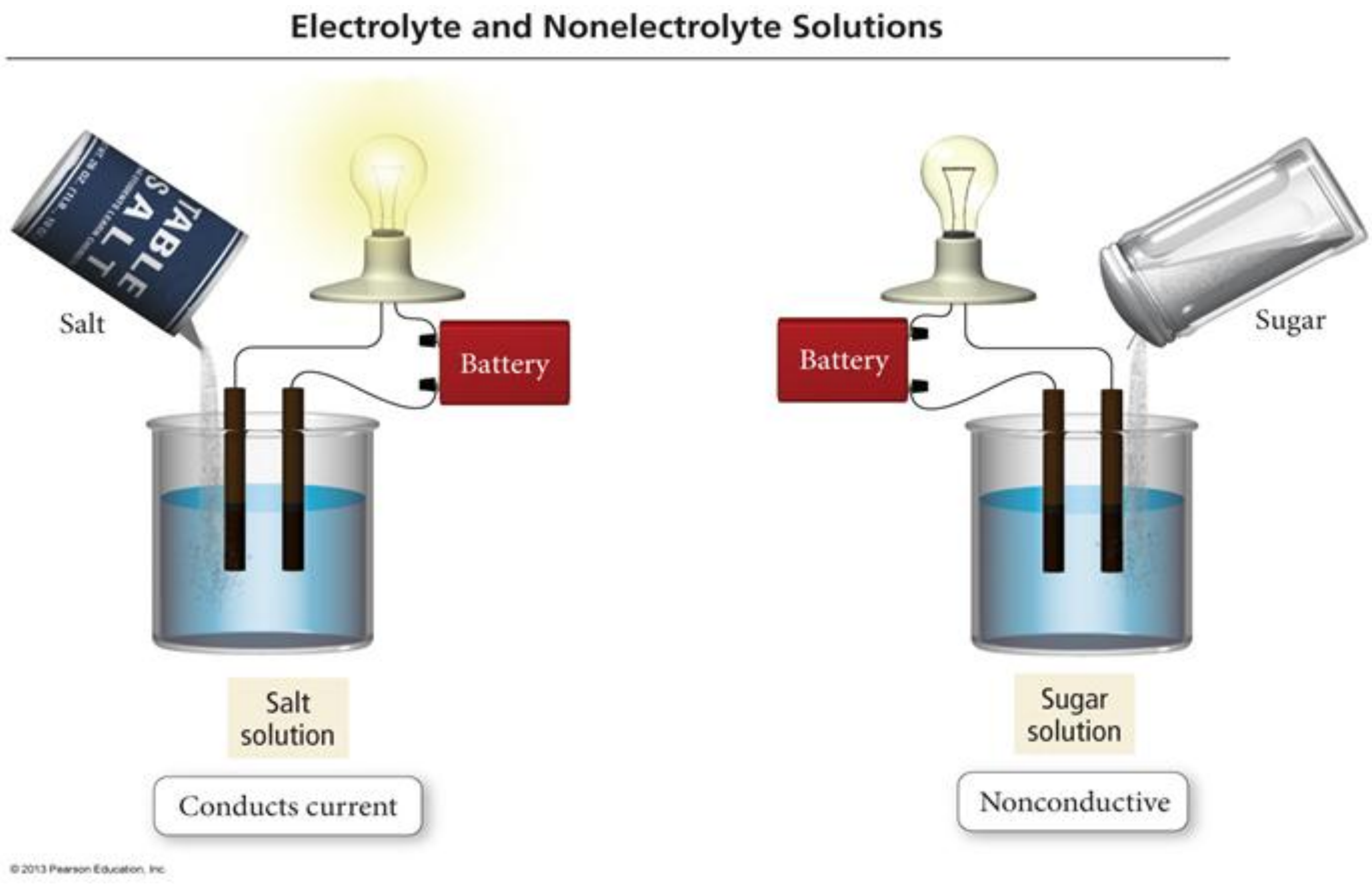
Dissolving Sugar in Water

- ✓ Sucrose ($C_{12}H_{22}O_{11}$) molecules homogeneously mixed with the water molecules.
- ✓ Compounds such as sugar that do not dissociate into ions when dissolved in water are called nonelectrolytes, and the resulting solutions do not conduct electricity.



Electrolytes and Nonelectrolytes

Substances that dissolve in water to form solutions that conduct electricity are called Electrolytes.

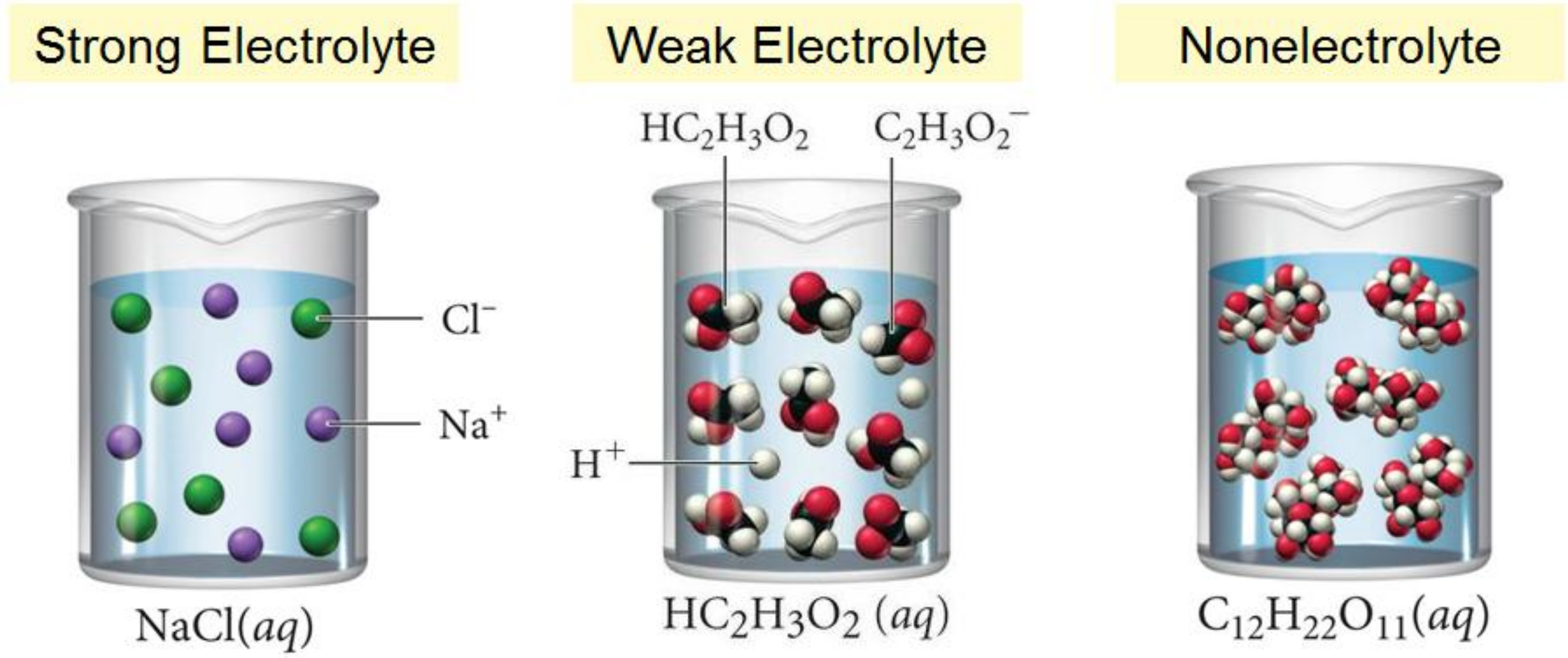


- **Solution of salt (an electrolyte)**
- **Solution of sugar (a nonelectrolyte)**

Electrolyte and Nonelectrolyte Solutions

- ➢ **Strong Electrolytes:**
  - Include substances that completely ionize when dissolve in water.
    - ✓ They can conduct electrical current strongly.
    - ✓ Examples: Soluble ionic salts ($NaCl$, $MgBr_2$ …), strong acids (HCl or HNO_3) or strong bases ($NaOH$ or $Mg(OH)_2$).
- ➢ **Weak Electrolytes:**
  - Include substances that partially ionize when dissolve in water.
    - ✓ They can conduct electrical current weakly.
    - ✓ Examples: weak acids (HF or CH_3COOH) or weak bases (NH_4OH).
- ➢ **Nonelectrolytes:**
  - Include substances that do not ionize when dissolve in water.
    - ✓ They don’t conduct electrical current.
    - ✓ Example: polar molecular substances (such as sugar or alcohol).

Electrolytes and Nonelectrolytes: A Summary



Complete Ionizing in water (full dissociation)
Examples: ionic salts, strong acids & strong bases

Partial Ionizing in water (partial dissociation)
Examples: weak acids & weak bases

No Ionizing in water (no dissociation)
Examples: many molecular (covalent) compounds as sugar

Assessment

All of the following compounds are soluble in water, indicate which of them is expected to produce strong, weak or non-electrolyte solution?

a. CsCl

b. CH_3OH

c. $Ca(NO_2)_2$

d. $C_6H_{12}O_6$

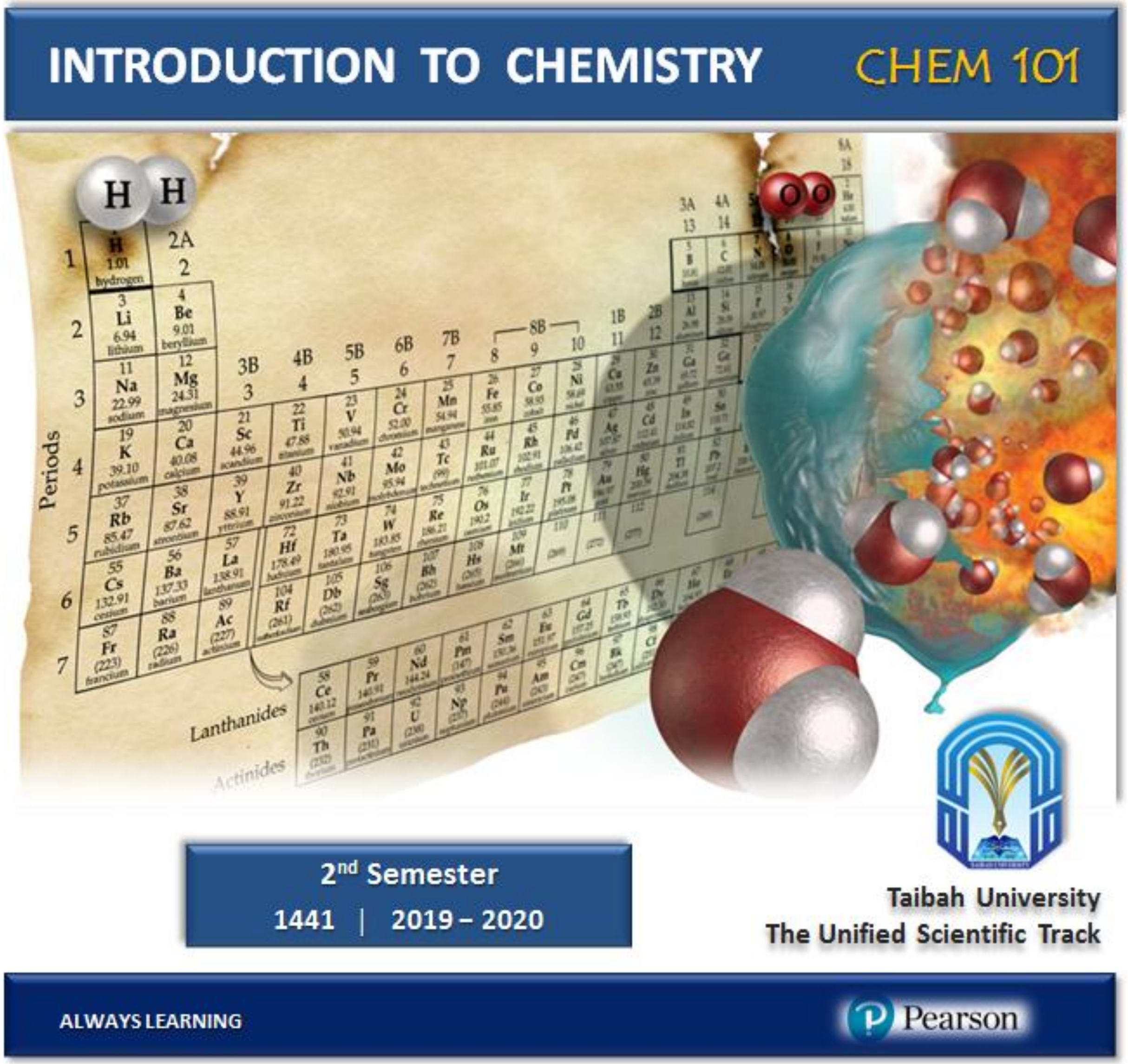
e. Acetic acid, vinegar (CH_3COOH) (weak acid)

f. HCl (strong acid)

g. NaOH (strong base)

h. HF (weak acid)

i. NH_4OH (weak base)



INTRODUCTION TO CHEMISTRY CHEM 101

2^{nd} Semester
1441 | 2019 – 2020

Taibah University
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ALWAYS LEARNING

Pearson

Lecture Presentation

Chapter 4

Stoichiometry, Solution Concentration and Chemical Reactions

Topic 15

- **Acid – Base Reactions**
- **Oxidation – Reduction Reactions**

Acids

- **Acid:** a substance that produces H^+ ions (*also known as H-protons*) in aqueous solutions:

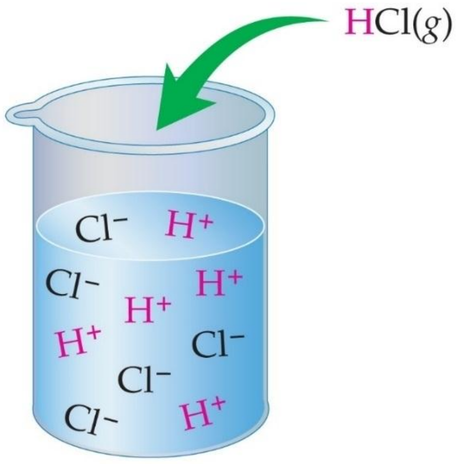
$$HCl(aq) \longrightarrow H^+(aq) + Cl^-(aq)$$

Examples (Important):

- **Strong Acids:** HCl, HBr, HI, H_2SO_4, HNO_3
- **Weak Acids:** acetic acid (vinegar, CH_3COOH), HF

Note: H^+ ion is a bare proton. Protons associate with water molecules in solution to form **hydronium ions (H_3O^+):**

$$H^+(aq) + H_2O(l) \rightarrow H_3O^+(aq)$$



HCl is an acid because it dissociates in water to produce $H^+(aq)$ ions.

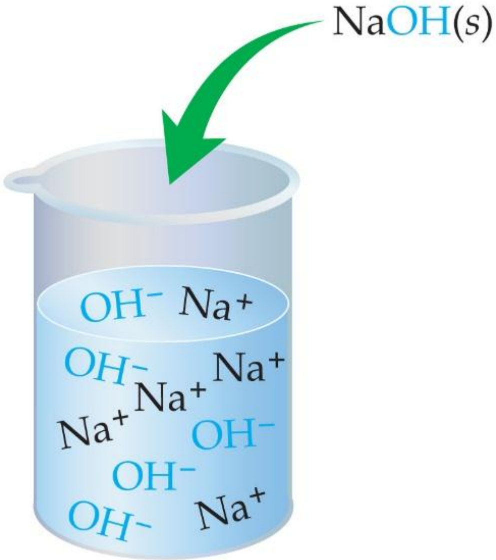
Bases (Alkalis)

- **Base (Also known as Alkali):** a substance that produces OH^- ions (*hydroxide ions*) in aqueous solutions:

$$NaOH(aq) \longrightarrow Na^+(aq) + OH^-(aq)$$

Examples (Important):

- **Strong Bases:** NaOH, KOH, LiOH, $Ba(OH)_2$ and $Ca(OH)_2$
- **Weak Base:** NH_4OH



NaOH is a base because it produces $OH^-(aq)$ ions when added to water.

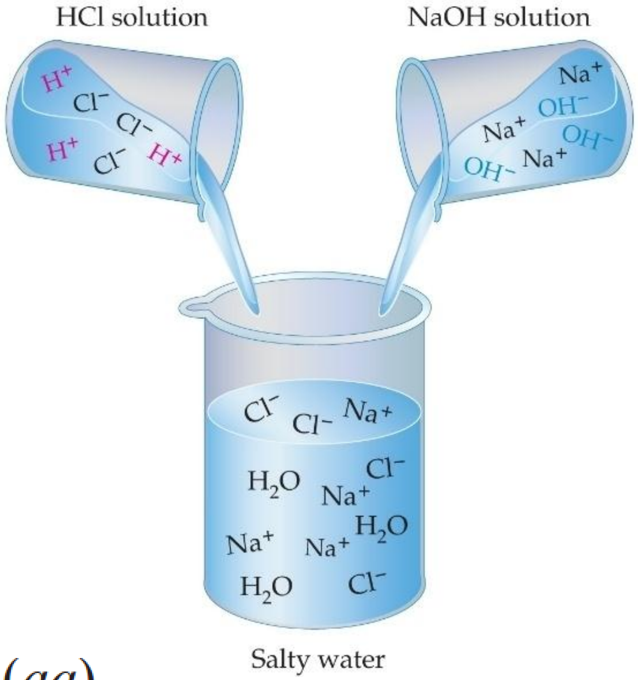
4.7 Acid – Base Reactions (Neutralization Reactions)

- **Acid–Base Neutralization Reactions:**

$$\boxed{\textbf{Acid} + \textbf{Base} \longrightarrow \textbf{Water} + \textbf{Salt} \ \text{(acid–base reactions)}}$$

- **Example:**

$$\underset{\text{Acid}}{HCl(aq)} + \underset{\text{Base}}{NaOH(aq)} \longrightarrow \underset{\text{Water}}{H_2O(l)} + \underset{\text{Salt}}{NaCl(aq)}$$



- **Net ionic equation:**

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

- **Another Example:**

$$\underset{\text{acid}}{HI(aq)} + \underset{\text{base}}{Ba(OH)_2(aq)} \longrightarrow \underset{\text{water}}{H_2O(l)} + \underset{\text{salt}}{BaI_2(aq)}$$

Oxidation – Reduction Reactions (Redox)

Oxidation–reduction reactions or redox reactions are reactions in which electrons are transferred from one reactant to the other.

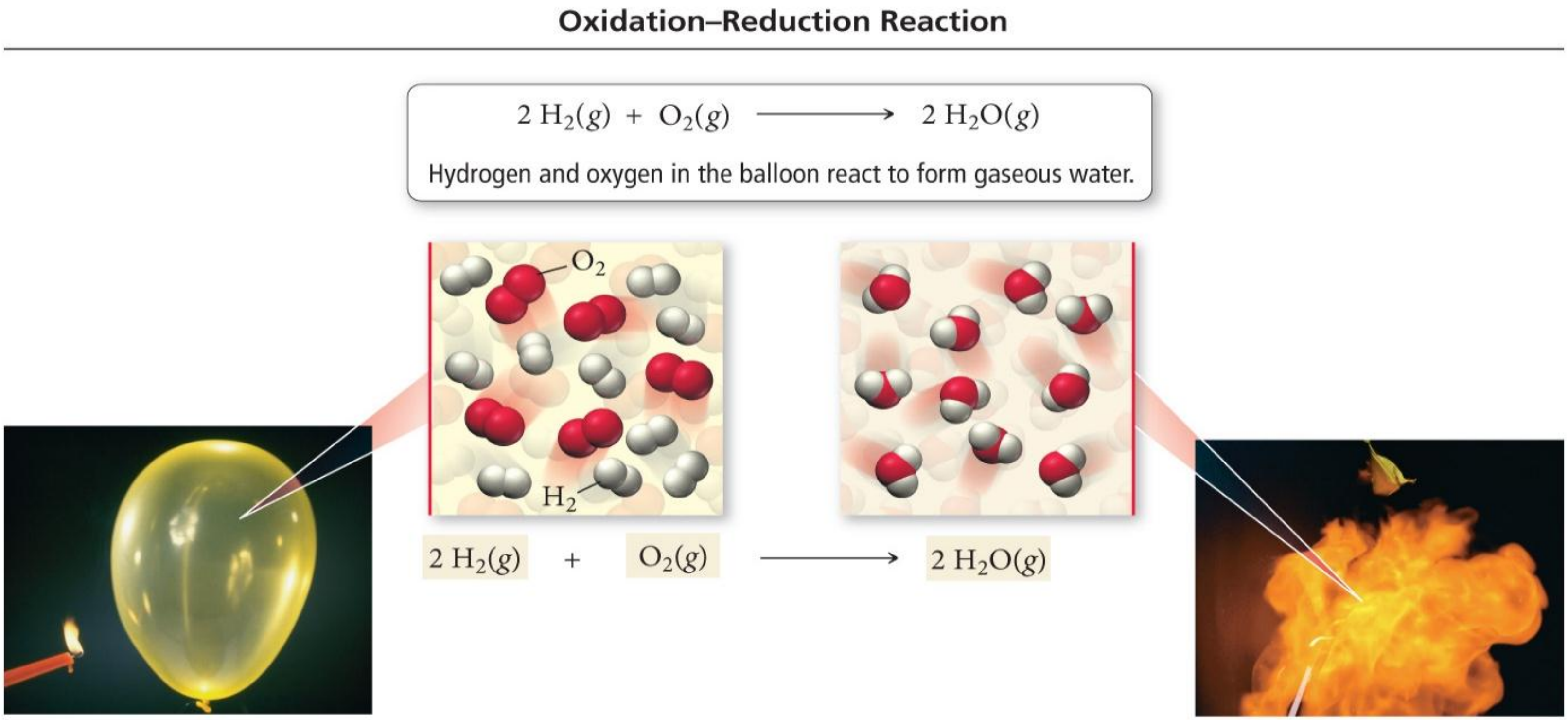
- **Oxidation:** is the **loss** of electrons.
- **Reduction:** is the **gain** of electrons.
  - ✓ Based on these definitions, redox reactions do not need to involve oxygen.
  - ✓ One cannot occur without the other.

➢ **Example on Redox Reactions:**

$$2\,\mathrm{Na}(s) + \mathrm{Cl_2}(g) \longrightarrow 2\,\mathrm{NaCl}(s)$$

✓ In this reaction, a metal (which has a tendency to lose electrons) reacts with a nonmetal (which has a tendency to gain electrons). In other words, metal atoms lose electrons to nonmetal atoms.

Oxidation – Reduction Reactions (Redox): Examples



Oxidation–Reduction Reaction

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

Hydrogen and oxygen in the balloon react to form gaseous water.

- **Other common redox reactions:**

$4\ Fe(s) + 3\ O_2(g) \longrightarrow 2\ Fe_2O_3(s)$ (rusting of iron)

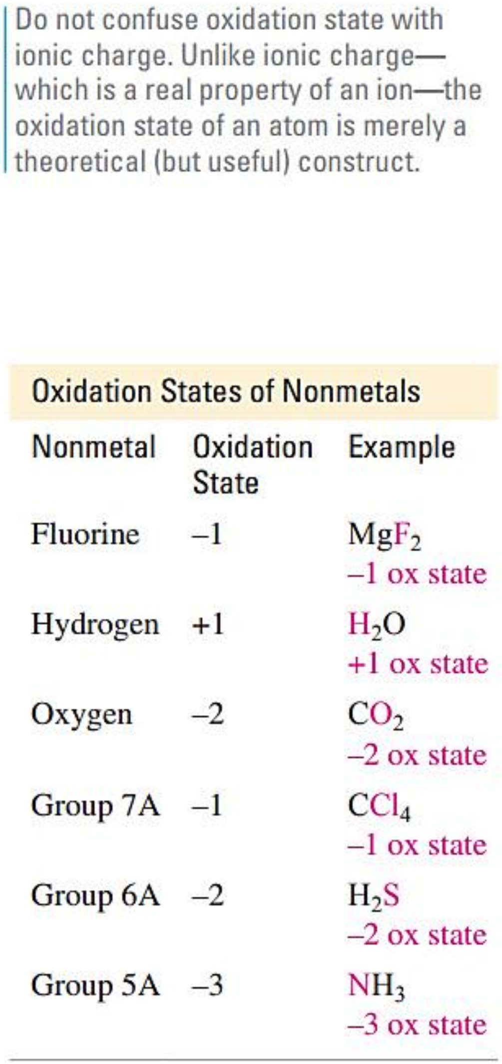
$2\ C_8H_{18}(l) + 25\ O_2(g) \longrightarrow 16\ CO_2(g) + 18\ H_2O(g)$ (combustion of octane)

$2\ H_2(g) + O_2(g) \longrightarrow 2\ H_2O(g)$ (combustion of hydrogen)

Oxidation Numbers (Oxidation States)

- ✓ Before we can identify an oxidation-reduction reaction, we must keep track of **lost** and **gained** electrons
- ✓ To do this, the concept of **oxidation numbers** (also called oxidation states) is used.
- ✓ Each atom in a neutral substance or ion is assigned an oxidation number.
- ✓ See the “Rules for Assigning Oxidation States” in the next section.

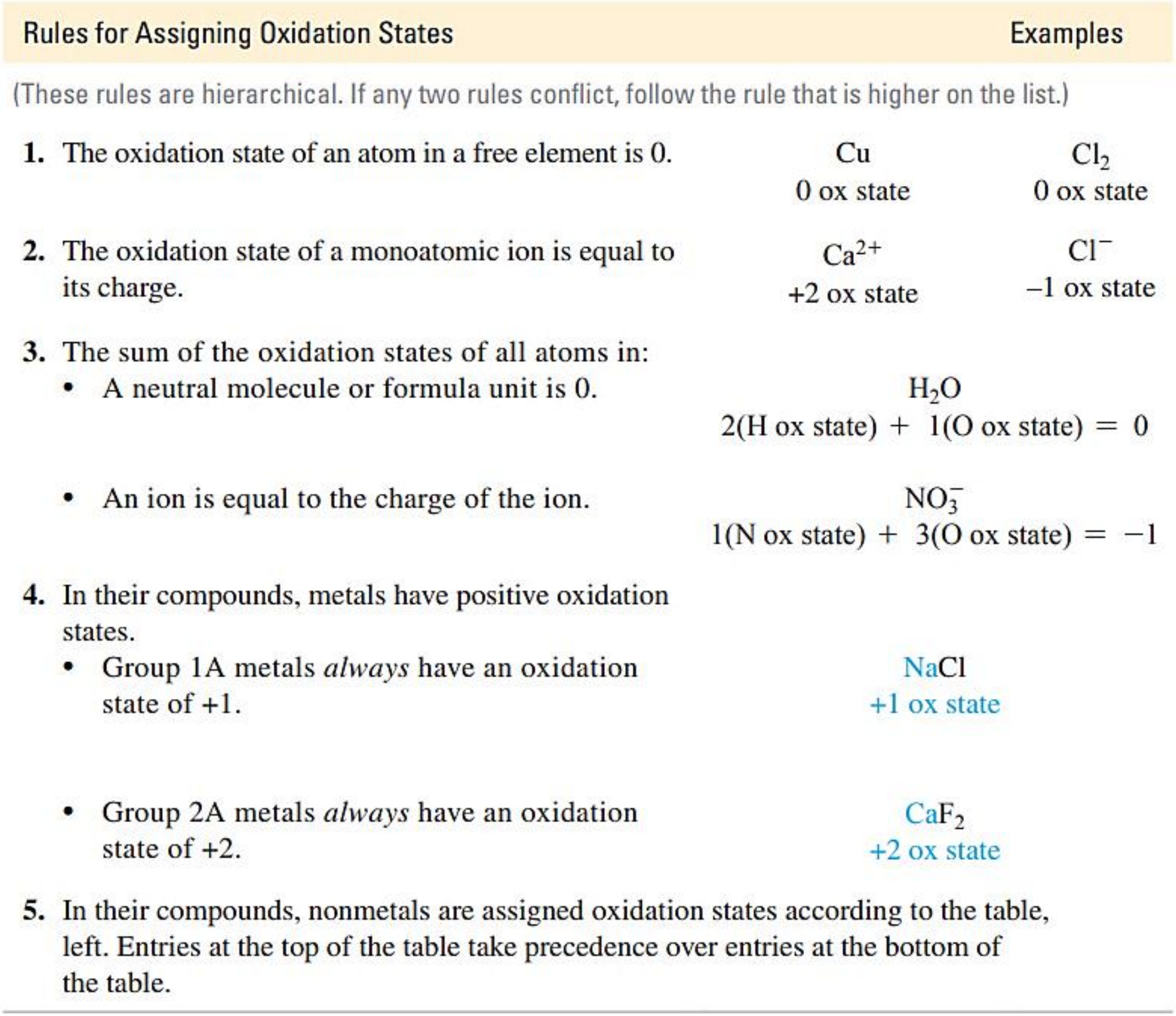
Rules for Assigning Oxidation States



Do not confuse oxidation state with ionic charge. Unlike ionic charge—which is a real property of an ion—the oxidation state of an atom is merely a theoretical (but useful) construct.

Oxidation States of Nonmetals

Nonmetal	Oxidation State	Example
Fluorine	–1	MgF_2 –1 ox state
Hydrogen	+1	H_2O +1 ox state
Oxygen	–2	CO_2 –2 ox state
Group 7A	–1	CCl_4 –1 ox state
Group 6A	–2	H_2S –2 ox state
Group 5A	–3	NH_3 –3 ox state



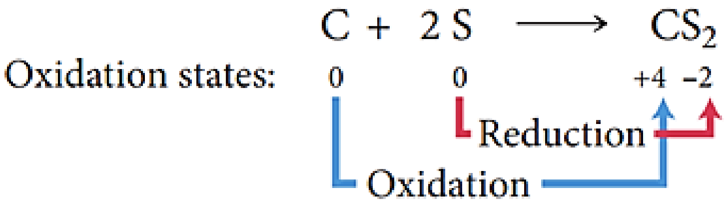
Rules for Assigning Oxidation States	Examples	
(These rules are hierarchical. If any two rules conflict, follow the rule that is higher on the list.)		
1. The oxidation state of an atom in a free element is 0.	Cu 0 ox state	Cl_2 0 ox state
2. The oxidation state of a monoatomic ion is equal to its charge.	Ca^{2+} +2 ox state	Cl^- –1 ox state
3. The sum of the oxidation states of all atoms in: • A neutral molecule or formula unit is 0.	H_2O 2(H ox state) + 1(O ox state) = 0	
• An ion is equal to the charge of the ion.	NO_3^- 1(N ox state) + 3(O ox state) = −1	
4. In their compounds, metals have positive oxidation states. • Group 1A metals *always* have an oxidation state of +1.	NaCl +1 ox state	
• Group 2A metals *always* have an oxidation state of +2.	CaF_2 +2 ox state	
5. In their compounds, nonmetals are assigned oxidation states according to the table, left. Entries at the top of the table take precedence over entries at the bottom of the table.		

Identifying Redox Reactions

Oxidation states can be used to identify redox reactions, even between nonmetals. For example, is the following reaction between carbon and sulfur a redox reaction?

$$C + 2\,S \longrightarrow CS_2$$

If so, what element is oxidized? What element is reduced? We can use the oxidation state rules to assign oxidation states to all elements on both sides of the equation.



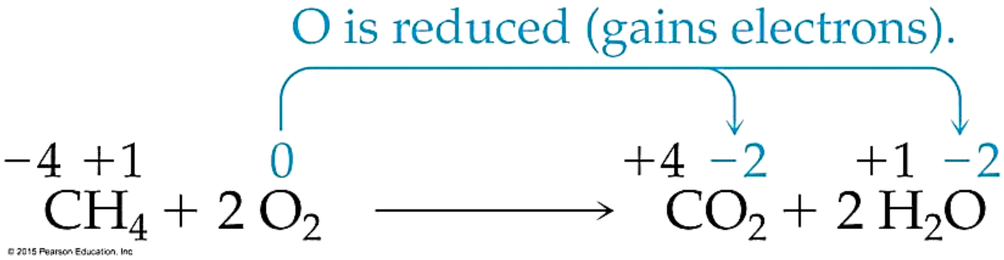
Carbon changed from an oxidation state of 0 to an oxidation state of +4. In terms of our electron bookkeeping scheme (the assigned oxidation state), carbon *lost electrons* and was *oxidized*. Sulfur changed from an oxidation state of 0 to an oxidation state of –2.

In terms of our electron bookkeeping scheme, sulfur *gained electrons* and was *reduced*. In terms of oxidation states, oxidation and reduction are defined as follows.

- **Oxidation An increase in oxidation state**
- **Reduction A decrease in oxidation state**

Identifying Redox Reactions

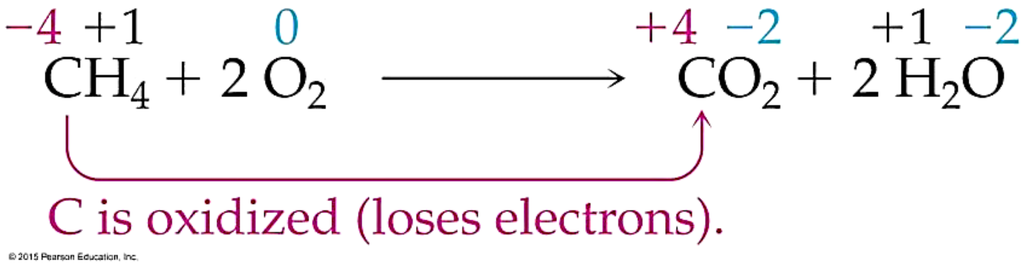
- **Reduction:**
  - The gaining of electrons, or:
  - Decrease in the oxidation state



O is reduced (gains electrons).

$$\overset{-4\ +1}{CH_4} + 2\,\overset{0}{O_2} \longrightarrow \overset{+4\ -2}{CO_2} + 2\,\overset{+1\ -2}{H_2O}$$

- **Oxidation:**
  - ✓ The loss of electrons, or:
  - ✓ Increase in the oxidation state



$$\overset{-4\ +1}{CH_4} + 2\,\overset{0}{O_2} \longrightarrow \overset{+4\ -2}{CO_2} + 2\,\overset{+1\ -2}{H_2O}$$

C is oxidized (loses electrons).

Identifying Redox Reactions: Practice

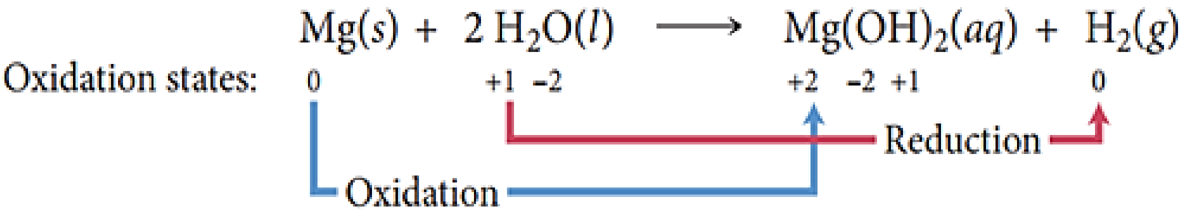
EXAMPLE 4.9 Using Oxidation States to Identify Oxidation and Reduction

Use oxidation states to identify the element that is being oxidized and the element that is being reduced in the following redox reaction.

$$Mg(s) + 2\,H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$$

Solution

Begin by assigning oxidation states to each atom in the reaction.



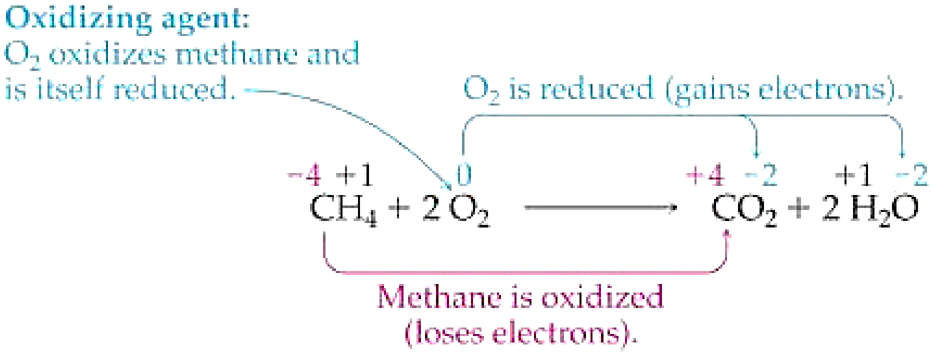
$$Mg(s) + 2\,H_2O(l) \longrightarrow Mg(OH)_2(aq) + H_2(g)$$

Oxidation states: 0 +1 −2 +2 −2 +1 0

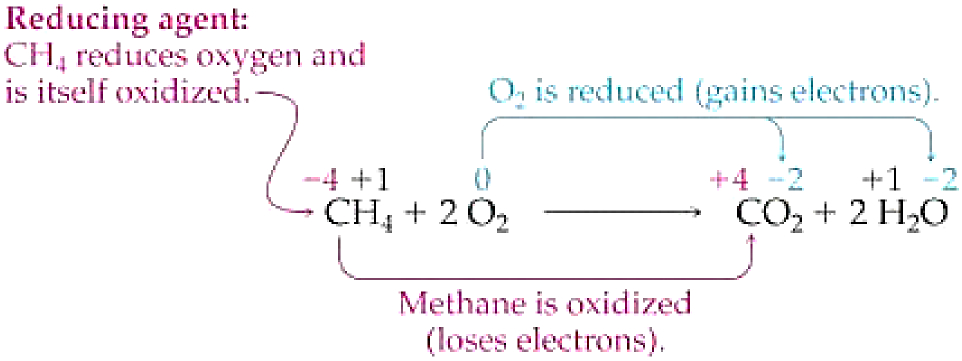
- ✓ Since Mg increased in oxidation state, it was oxidized.
- ✓ Since H decreased in oxidation state, it was reduced.

Oxidizing Agent & Reducing Agent

- **Oxidizing Agent:**
  - ✓ A Substance that **oxidizes** something else. The **oxidizing agent** itself is **reduced** in the same reaction:



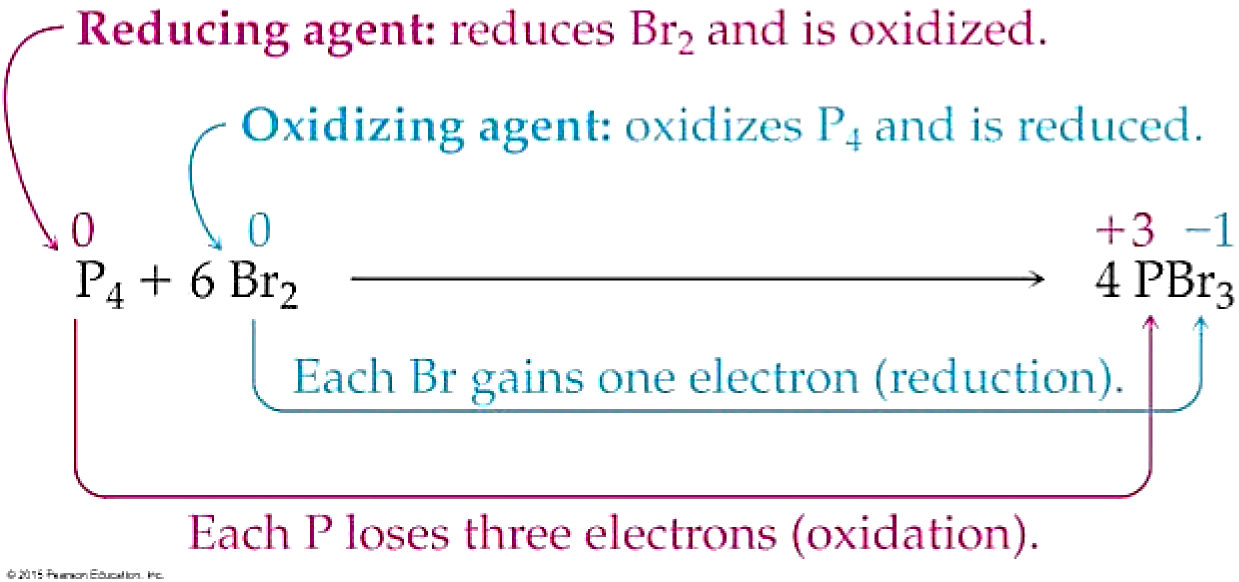
- **Reducing Agent:**
  - ✓ A Substance that **reduces** something else. The **reducing agent** itself is **oxidized** in the same reaction:



Oxidizing Agent & Reducing Agent: Practice

In the following reaction, identify the oxidizing agent and the reducing agent?

$$P_4(s) + 6\ Br_2(l) \rightarrow 4\ PBr_3(g)$$



Reducing agent: reduces Br_2 and is oxidized.

Oxidizing agent: oxidizes P_4 and is reduced.

$$\overset{0}{P_4} + 6\ \overset{0}{Br_2} \longrightarrow 4\ \overset{+3}{P}\overset{-1}{Br_3}$$

Each Br gains one electron (reduction).

Each P loses three electrons (oxidation).

Oxidizing Agent & Reducing Agent: Practice

Identifying Redox Reactions, Oxidizing Agents, and Reducing Agents

Determine whether each reaction is an oxidation–reduction reaction. If the reaction is an oxidation–reduction, identify the oxidizing agent and the reducing agent.

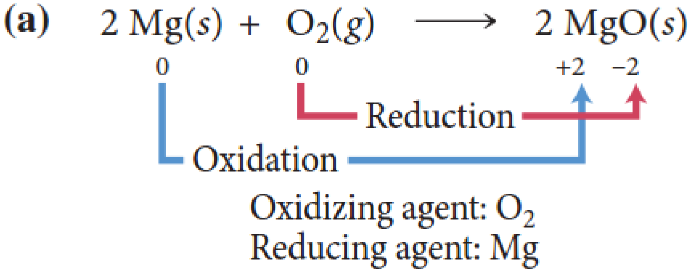
(a) $2\,Mg(s) + O_2(g) \longrightarrow 2\,MgO(s)$

(b) $2\,HBr(aq) + Ca(OH)_2(aq) \longrightarrow 2\,H_2O(l) + CaBr_2(aq)$

(c) $Zn(s) + Fe^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Fe(s)$

SOLUTION

This is a redox reaction because magnesium increases in oxidation number (oxidation) and oxygen decreases in oxidation number (reduction).



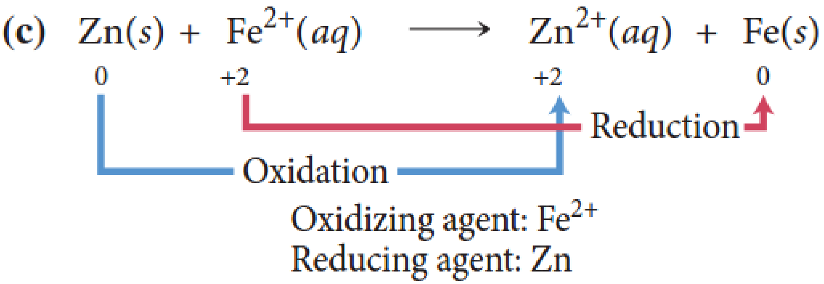
(a) $2\,\underset{0}{Mg}(s) + \underset{0}{O_2}(g) \longrightarrow 2\,\underset{+2\;\;-2}{MgO}(s)$

Oxidizing agent: O_2
Reducing agent: Mg

This is not a redox reaction because none of the atoms undergoes a change in oxidation number.

(b) $2\,\underset{+1\,-1}{HBr}(aq) + \underset{+2\;-2+1}{Ca(OH)_2}(aq) \rightarrow 2\underset{+1\,-2}{H_2O}(l) + \underset{+2\;-1}{CaBr_2}(aq)$

This is a redox reaction because zinc increases in oxidation number (oxidation) and iron decreases in oxidation number (reduction).



(c) $\underset{0}{Zn}(s) + \underset{+2}{Fe^{2+}}(aq) \longrightarrow \underset{+2}{Zn^{2+}}(aq) + \underset{0}{Fe}(s)$

Oxidizing agent: Fe^{2+}
Reducing agent: Zn

Assessment

1- Assign oxidation states to each atom in each ion or compound.

a. Ag **b.** Ag^+

c. CaF_2 **d.** H_2S

e. CO_3^{2-} **f.** CrO_4^{2-}

2- What is the oxidation state of Cr in each compound?

a. CrO

b. CrO_3

c. Cr_2O_3

3- Which reactions are redox reactions? For each redox reaction, identify the oxidizing agent and the reducing agent.

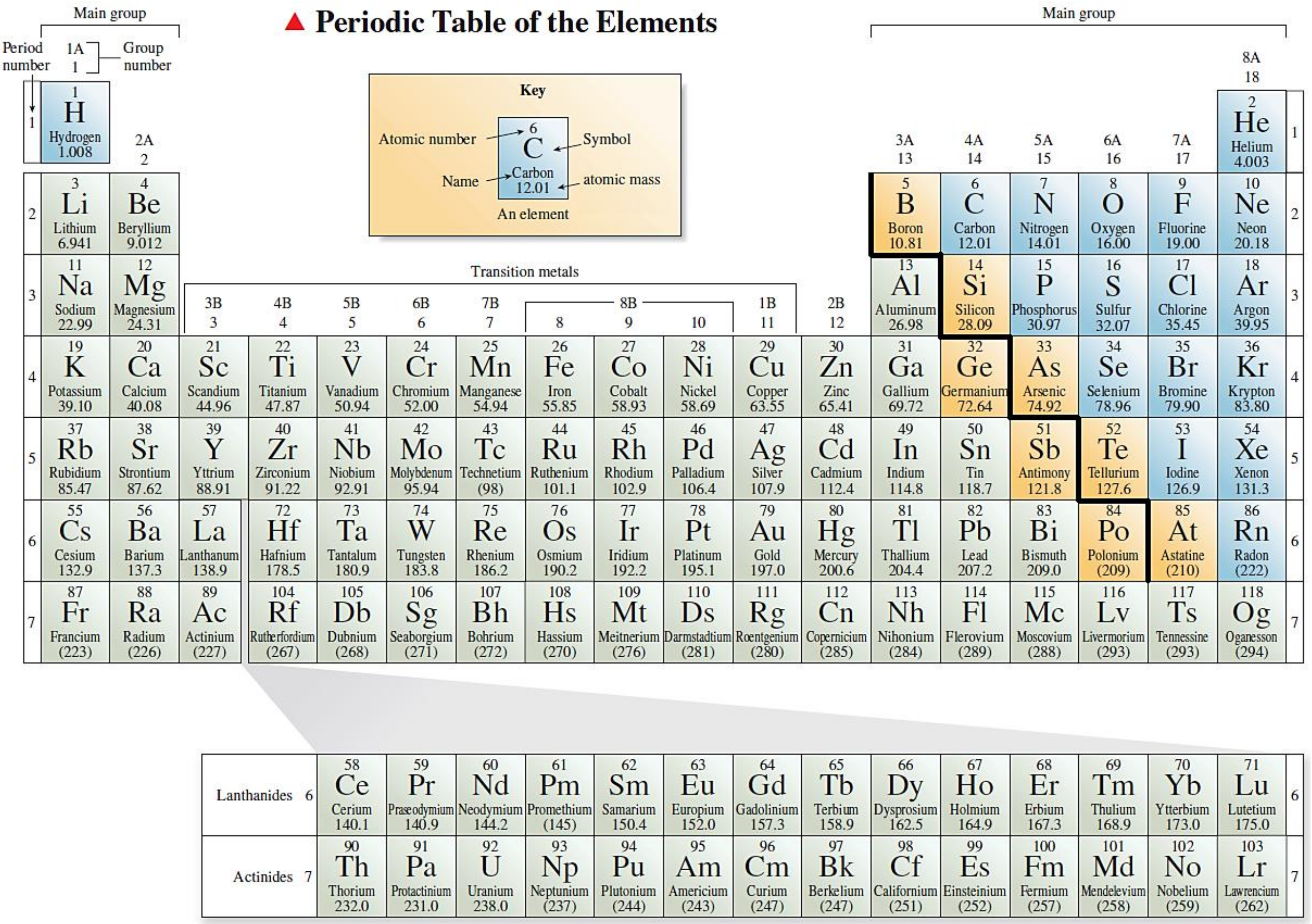
a. $4\,Li(s) + O_2(g) \rightarrow 2\,Li_2O(s)$

b. $Mg(s) + Fe^{2+}(aq) \rightarrow Mg^{2+}(aq) + Fe(s)$

c. $Pb(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow PbSO_4(s) + 2\,NaNO_3(aq)$

d. $HBr(aq) + KOH(aq) \rightarrow H_2O(l) + KBr(aq)$

▲ Periodic Table of the Elements



Period	1A	2A	3B	4B	5B	6B	7B	8B	8B	8B	1B	2B	3A	4A	5A	6A	7A	8A
	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H Hydrogen 1.008																	2 He Helium 4.003
2	3 Li Lithium 6.941	4 Be Beryllium 9.012											5 B Boron 10.81	6 C Carbon 12.01	7 N Nitrogen 14.01	8 O Oxygen 16.00	9 F Fluorine 19.00	10 Ne Neon 20.18
3	11 Na Sodium 22.99	12 Mg Magnesium 24.31											13 Al Aluminum 26.98	14 Si Silicon 28.09	15 P Phosphorus 30.97	16 S Sulfur 32.07	17 Cl Chlorine 35.45	18 Ar Argon 39.95
4	19 K Potassium 39.10	20 Ca Calcium 40.08	21 Sc Scandium 44.96	22 Ti Titanium 47.87	23 V Vanadium 50.94	24 Cr Chromium 52.00	25 Mn Manganese 54.94	26 Fe Iron 55.85	27 Co Cobalt 58.93	28 Ni Nickel 58.69	29 Cu Copper 63.55	30 Zn Zinc 65.41	31 Ga Gallium 69.72	32 Ge Germanium 72.64	33 As Arsenic 74.92	34 Se Selenium 78.96	35 Br Bromine 79.90	36 Kr Krypton 83.80
5	37 Rb Rubidium 85.47	38 Sr Strontium 87.62	39 Y Yttrium 88.91	40 Zr Zirconium 91.22	41 Nb Niobium 92.91	42 Mo Molybdenum 95.94	43 Tc Technetium (98)	44 Ru Ruthenium 101.1	45 Rh Rhodium 102.9	46 Pd Palladium 106.4	47 Ag Silver 107.9	48 Cd Cadmium 112.4	49 In Indium 114.8	50 Sn Tin 118.7	51 Sb Antimony 121.8	52 Te Tellurium 127.6	53 I Iodine 126.9	54 Xe Xenon 131.3
6	55 Cs Cesium 132.9	56 Ba Barium 137.3	57 La Lanthanum 138.9	72 Hf Hafnium 178.5	73 Ta Tantalum 180.9	74 W Tungsten 183.8	75 Re Rhenium 186.2	76 Os Osmium 190.2	77 Ir Iridium 192.2	78 Pt Platinum 195.1	79 Au Gold 197.0	80 Hg Mercury 200.6	81 Tl Thallium 204.4	82 Pb Lead 207.2	83 Bi Bismuth 209.0	84 Po Polonium (209)	85 At Astatine (210)	86 Rn Radon (222)
7	87 Fr Francium (223)	88 Ra Radium (226)	89 Ac Actinium (227)	104 Rf Rutherfordium (267)	105 Db Dubnium (268)	106 Sg Seaborgium (271)	107 Bh Bohrium (272)	108 Hs Hassium (270)	109 Mt Meitnerium (276)	110 Ds Darmstadtium (281)	111 Rg Roentgenium (280)	112 Cn Copernicium (285)	113 Nh Nihonium (284)	114 Fl Flerovium (289)	115 Mc Moscovium (288)	116 Lv Livermorium (293)	117 Ts Tennessine (293)	118 Og Oganesson (294)

Lanthanides 6	58 Ce Cerium 140.1	59 Pr Praseodymium 140.9	60 Nd Neodymium 144.2	61 Pm Promethium (145)	62 Sm Samarium 150.4	63 Eu Europium 152.0	64 Gd Gadolinium 157.3	65 Tb Terbium 158.9	66 Dy Dysprosium 162.5	67 Ho Holmium 164.9	68 Er Erbium 167.3	69 Tm Thulium 168.9	70 Yb Ytterbium 173.0	71 Lu Lutetium 175.0
Actinides 7	90 Th Thorium 232.0	91 Pa Protactinium 231.0	92 U Uranium 238.0	93 Np Neptunium (237)	94 Pu Plutonium (244)	95 Am Americium (243)	96 Cm Curium (247)	97 Bk Berkelium (247)	98 Cf Californium (251)	99 Es Einsteinium (252)	100 Fm Fermium (257)	101 Md Mendelevium (258)	102 No Nobelium (259)	103 Lr Lawrencium (262)