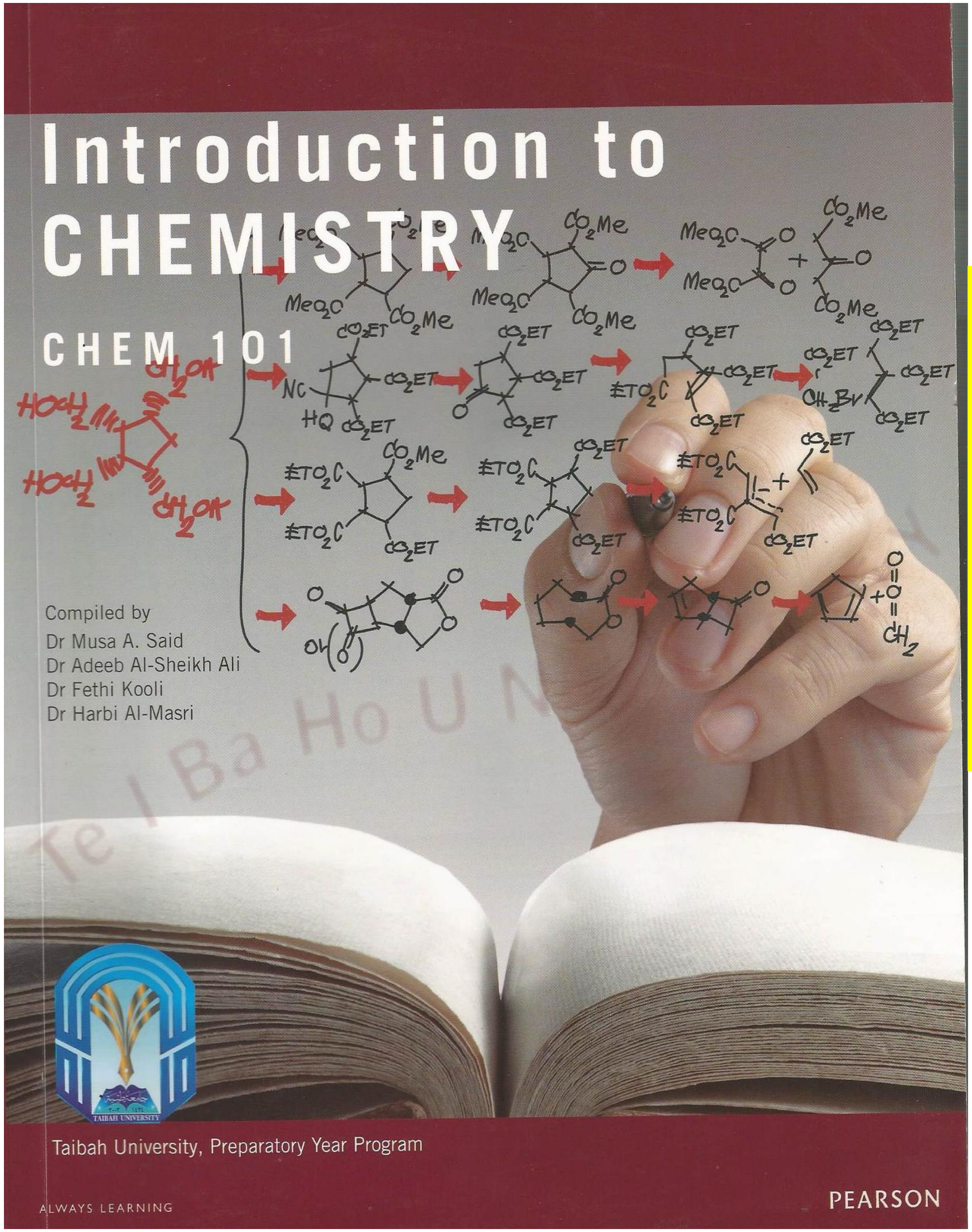

Introduction to
CHEMISTRY

CHEM 101

Compiled by
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Taibah University, Preparatory Year Program

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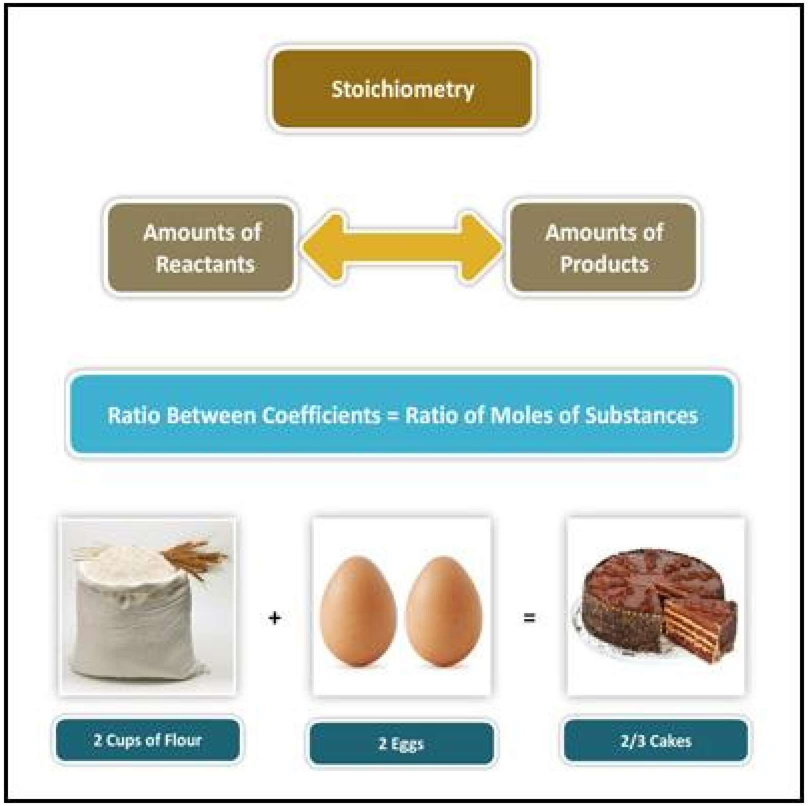
Lecture Presentation

Chapter 4

CHEMICAL BONDING
CHEMICAL REACTION

TOPIC 12
REACTION STOICHIOMETRY, LIMITING REACTANTS AND PRESENT YIELD

Introduction



OBJECTIVES

4.1.1 Identify stoichiometric quantities
4.1.2 Calculate limiting reactants, theoretical and percent yield

Just as the spe-cific prod-uct of a chem-i-cal re-ac-tion is im-por-tant, so is the quan-tity that can be pro-duced from the re-ac-tants and the ef-fi-ciency of that re-ac-tion.

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Stoichiometry

1. **Calculation of the quantities of reactants and products in a chemical reaction.**

2. **The quantitative relationship between reactants and products in a chemical reaction.**

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4.1 Reaction Stoichiometry: How much carbon dioxide?

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

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

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

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

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Making Molecules: Mole-to-Mole Conversion

From the balanced equation for the combustion of octane, we can write the following stoichiometric ratio:

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

Conversion factor

Suppose that we burn 22 moles of **C_8H_{18}** the amount of **CO_2** produced will be according to this conversion factor :

22 moles C_8H_{18} X16 moles CO_2 / 2 molesC_8H_{18}

= 176 moles CO_2

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Making Molecules: Mass-to-Mass Conversion

Mass A → Amount A (in moles) → Amount B (in moles) → Mass B

gC_8H_{18} → mol C_8H_{18} → mol CO_2 → gCO_2

1moleC_8H_{18} /
114.22g C_8H_{18}

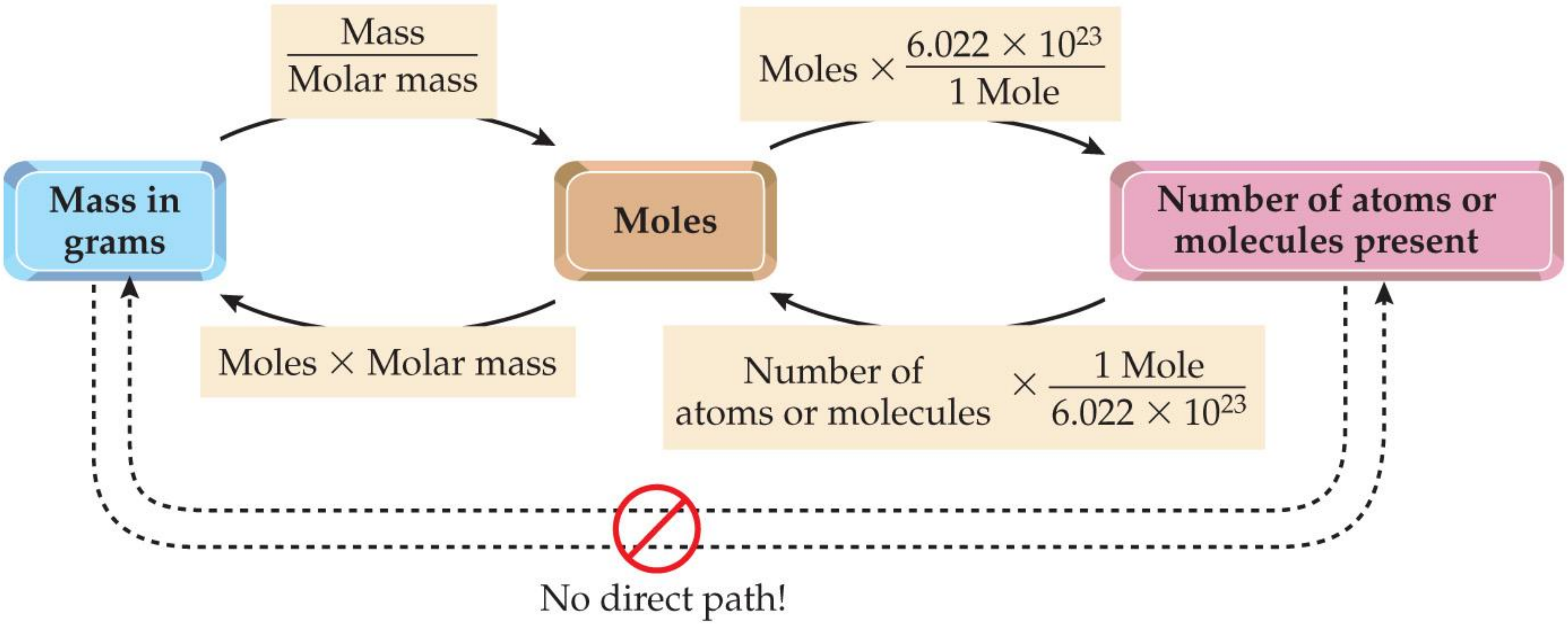
16moleCO_2 /
2mole C_8H_{18}

44gCO_2 /
1mole CO_2

Estimate the amount of CO_2 emitted into the atmosphere from the burning of $3.5x10^{15}$ g octane

Stoichiometry Conversions

anim_Stoichiometry_19_01.mp4



Conversions in Stoichiometry Calculations

Mass in grams

$\frac{\text{Mass}}{\text{Molar mass}}$

Moles

$\text{Moles} \times \frac{6.022 \times 10^{23}}{1\ \text{Mole}}$

Number of atoms or molecules present

$\text{Number of atoms or molecules} \times \frac{1\ \text{Mole}}{6.022 \times 10^{23}}$

$\text{Moles} \times \text{Molar mass}$

No direct path!

Example: Estimate the mass of CO_2 produced in 2007 by the combustion of 3.5×10^{15} g gasoline

Given: 3.4×10^{15} g C_8H_{18}

Find: g CO_2

Conceptual Plan:

$$\text{g } C_8H_{18} \xrightarrow{\frac{1\text{ mol}}{114.22\text{ g}}} \text{mol } C_8H_{18} \xrightarrow{\frac{16\text{ mol } CO_2}{2\text{ mol } C_8H_{18}}} \text{mol } CO_2 \xrightarrow{\frac{44.01\text{ g}}{1\text{ mol}}} \text{g } CO_2$$

Relationships: 1 mol C_8H_{18} = 114.22g, 1 mol CO_2 = 44.01g, 2 mol C_8H_{18}:16 mol CO_2

Solution:

$$3.5\times10^{15}\ \cancel{\text{g } C_8H_{18}} \times \frac{1\ \cancel{\text{mol } C_8H_{18}}}{114.22\ \cancel{\text{g } C_8H_{18}}} \times \frac{16\ \cancel{\text{mol } CO_2}}{2\ \cancel{\text{mol } C_8H_{18}}} \times \frac{44.01\text{ g } CO_2}{1\ \cancel{\text{mol } CO_2}}$$

$$= 1.1\times10^{16}\text{ g } CO_2$$

Check: because 8x moles of CO_2 as C_8H_{18}, but the molar mass of C_8H_{18} is 3x CO_2, the number makes sense

Example 4.1: How many grams of glucose can be synthesized from 37.8 g of CO_2 in photosynthesis?

Given: 37.8 g CO_2, $6\,CO_2 + 6\,H_2O \rightarrow C_6H_{12}O_6 + 6\,O_2$

Find: g $C_6H_{12}O_6$

Conceptual Plan:

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

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

Solution:

$$37.8\,\cancel{\text{g } CO_2} \times \frac{1\,\cancel{\text{mol } CO_2}}{44.01\,\cancel{\text{g } CO_2}} \times \frac{1\,\cancel{\text{mol } C_6H_{12}O_6}}{6\,\cancel{\text{mol } CO_2}} \times \frac{180.2\text{ g } C_6H_{12}O_6}{1\,\cancel{\text{mol } C_6H_{12}O_6}}$$

$$= 25.8\text{ g } C_6H_{12}O_6$$

Check: because 6x moles of CO_2 as $C_6H_{12}O_6$, but the molar mass of $C_6H_{12}O_6$ is 4x CO_2, the number makes sense

4.2 Limiting Reactant, Theoretical Yield, and Percent Yield

- Recall our pizza recipe: وصفة عمل البيتزا

1 crust + 5 oz tomato sauce + 2 cups cheese → 1 pizza

1 قشرة + 5 اوقية صلصة طماطم + 2 كوب جبن → 1 بيتزا

- If we have 4 crusts, 10 cups of cheese, and 15 oz tomato sauce. How many pizzas can we n

$$4\ \cancel{\text{crusts}} \times \frac{1\ \text{pizza}}{1\ \cancel{\text{crust}}} = 4\ \text{pizzas}$$

We have enough crusts to make

$$10\ \cancel{\text{cups cheese}} \times \frac{1\ \text{pizza}}{2\ \cancel{\text{cups cheese}}} = 5\ \text{pizzas}$$

We have enough cheese to make

We have enough tomato sauce to make

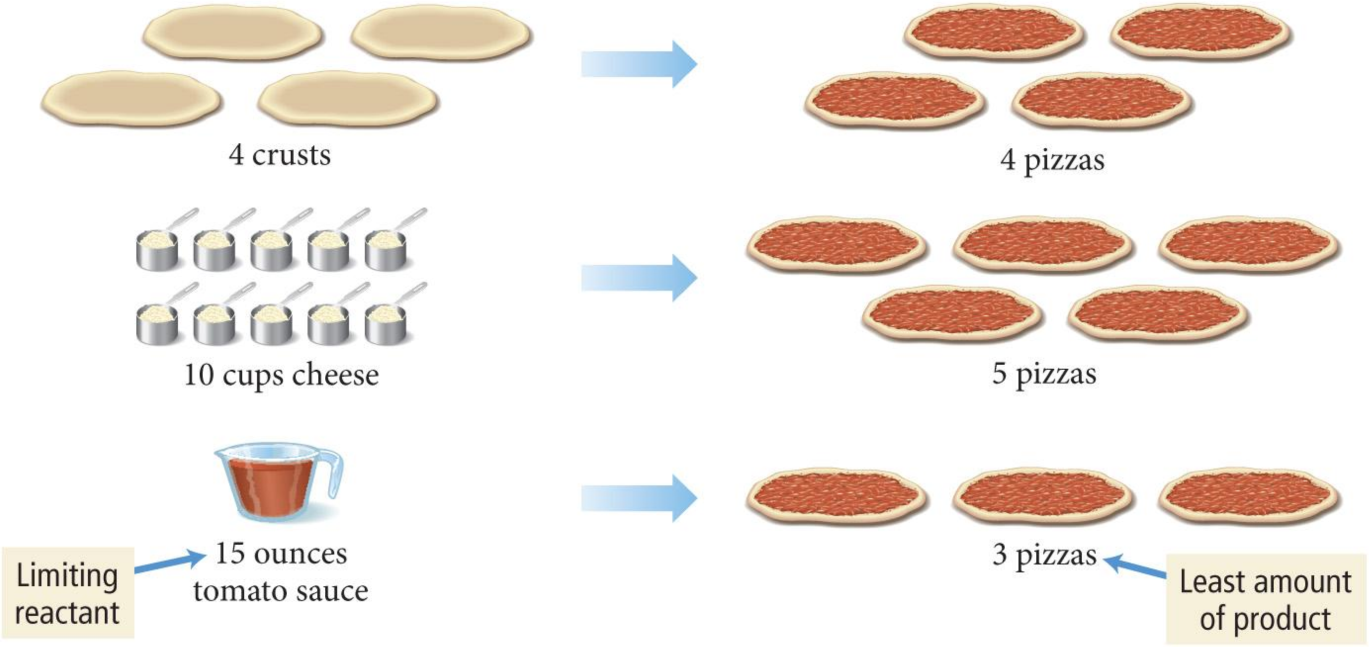
$$15\ \cancel{\text{ounces tomato sauce}} \times \frac{1\ \text{pizza}}{5\ \cancel{\text{ounces tomato sauce}}} = 3\ \text{pizzas}$$

Limiting reactant

Smallest number of pizzas

Limiting Reactant

- We have enough crusts for four pizzas, enough cheese for five pizzas, but enough tomato sauce for only three pizzas.
  - ✓ We can make only three pizzas. The tomato sauce *limits* how many pizzas we can make.



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Summarizing Limiting Reactant and Yield

- **The theoretical yield** is the amount of product that can be made in a chemical reaction based on the amount of limiting reactant.
- **The actual yield** is the amount of product actually produced by a chemical reaction.
- **The percent yield** is calculated as follows:

$$\frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

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Reaction Stoichiometry

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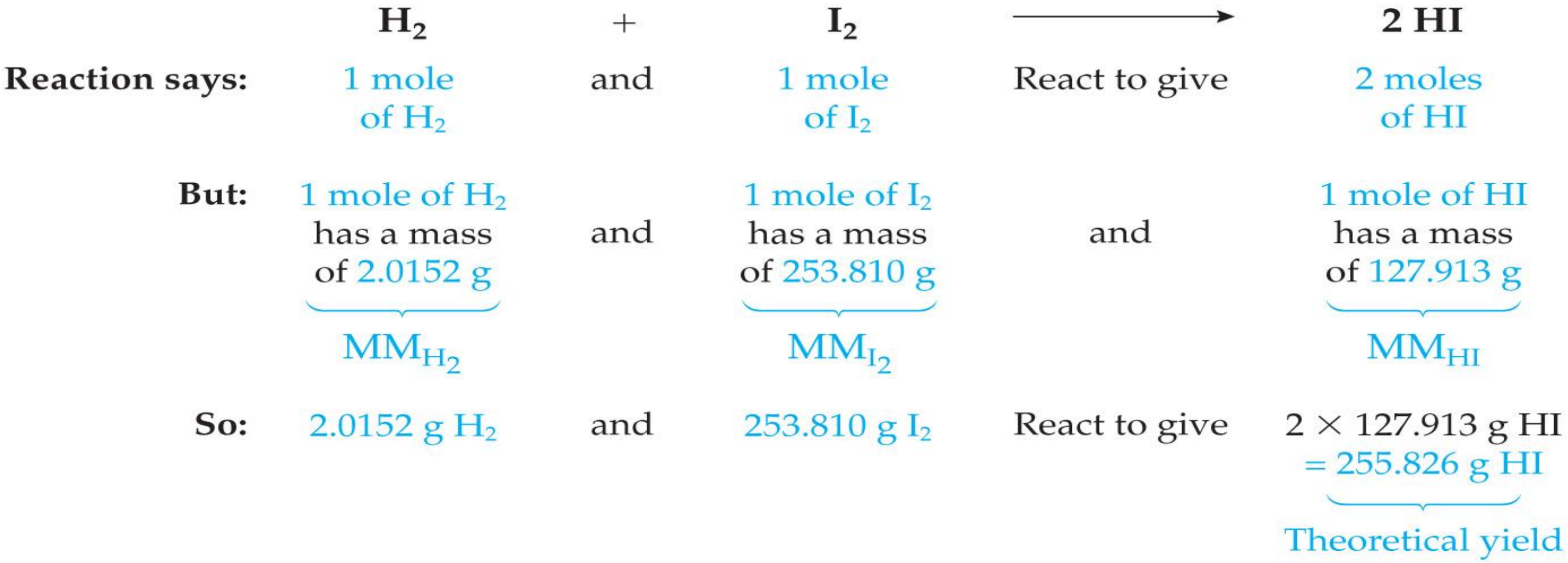
- **Theoretical yield**
  - ✓ Maximum amount of product that can be produced from a certain reactant amount



	H_2	+	I_2	$\longrightarrow$	2 HI
Reaction says:	1 mole of H_2	and	1 mole of I_2	React to give	2 moles of HI
But:	1 mole of H_2 has a mass of 2.0152 g MM_{H_2}	and	1 mole of I_2 has a mass of 253.810 g MM_{I_2}	and	1 mole of HI has a mass of 127.913 g MM_{HI}
So:	2.0152 g H_2	and	253.810 g I_2	React to give	2×127.913 g HI = 255.826 g HI Theoretical yield

Reaction Stoichiometry

- **Actual yield**
  - ✓ Amount actually produced in a chemical reaction
  - ✓ Less than theoretical yield because of side reactions and side products
  - ✓ Conservation of mass is still observed
- **Limiting reactant**
  - ✓ Reactant that is in short supply
  - ✓ Limits the amount of product that can be made
  - ✓ Will be completely used up in a reaction
- **Excess reactant**
  - ✓ Present in excess in a reaction
  - ✓ Will be left over after the reaction

Reaction Stoichiometry

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

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

The **percent yield** would be 87.990%.

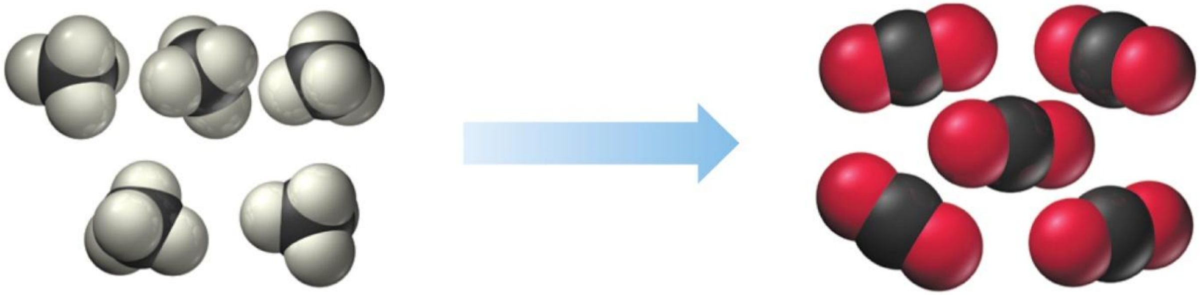
Apply These Concepts to a Chemical Reaction

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

$$CH_4(g) + \mathbf{2}\ O_2(g) \rightarrow CO_2(g) + \mathbf{2}\ H_2O(g)$$

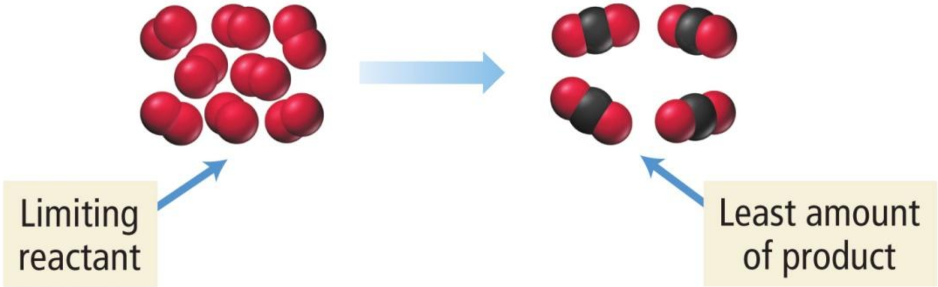
✓ First we calculate the number of CO_2 molecules that can be made from 5 CH_4 molecules.

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



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

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



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

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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.

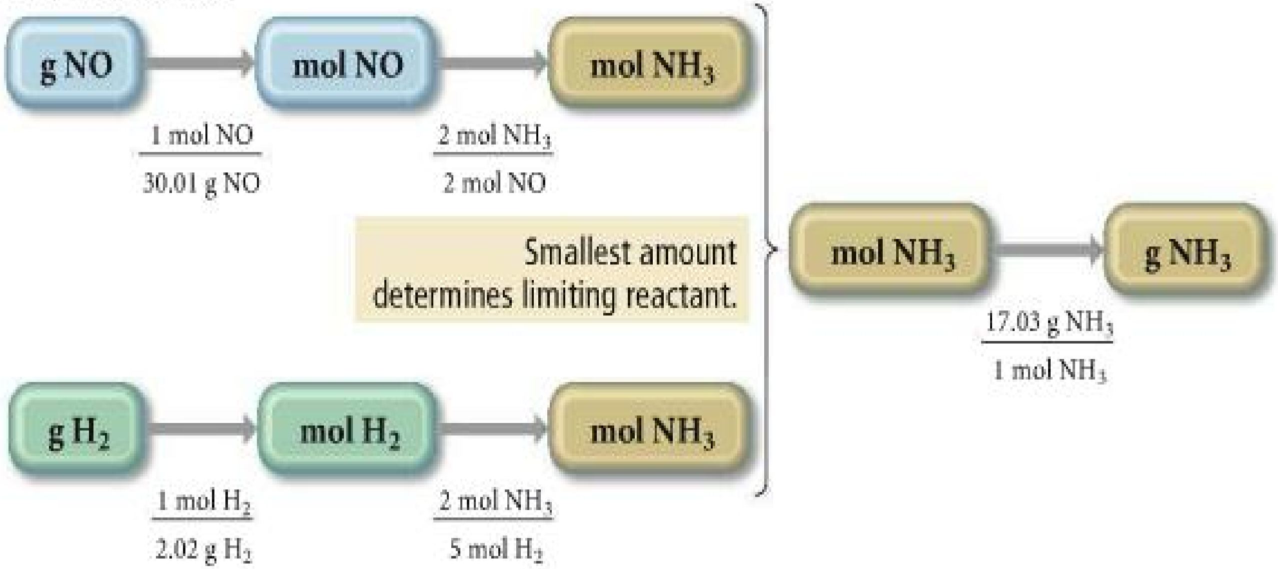
SORT You are given the mass of each reactant in grams and asked to find the theoretical yield of a product.

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 between moles of reactant

CONCEPTUAL PLAN

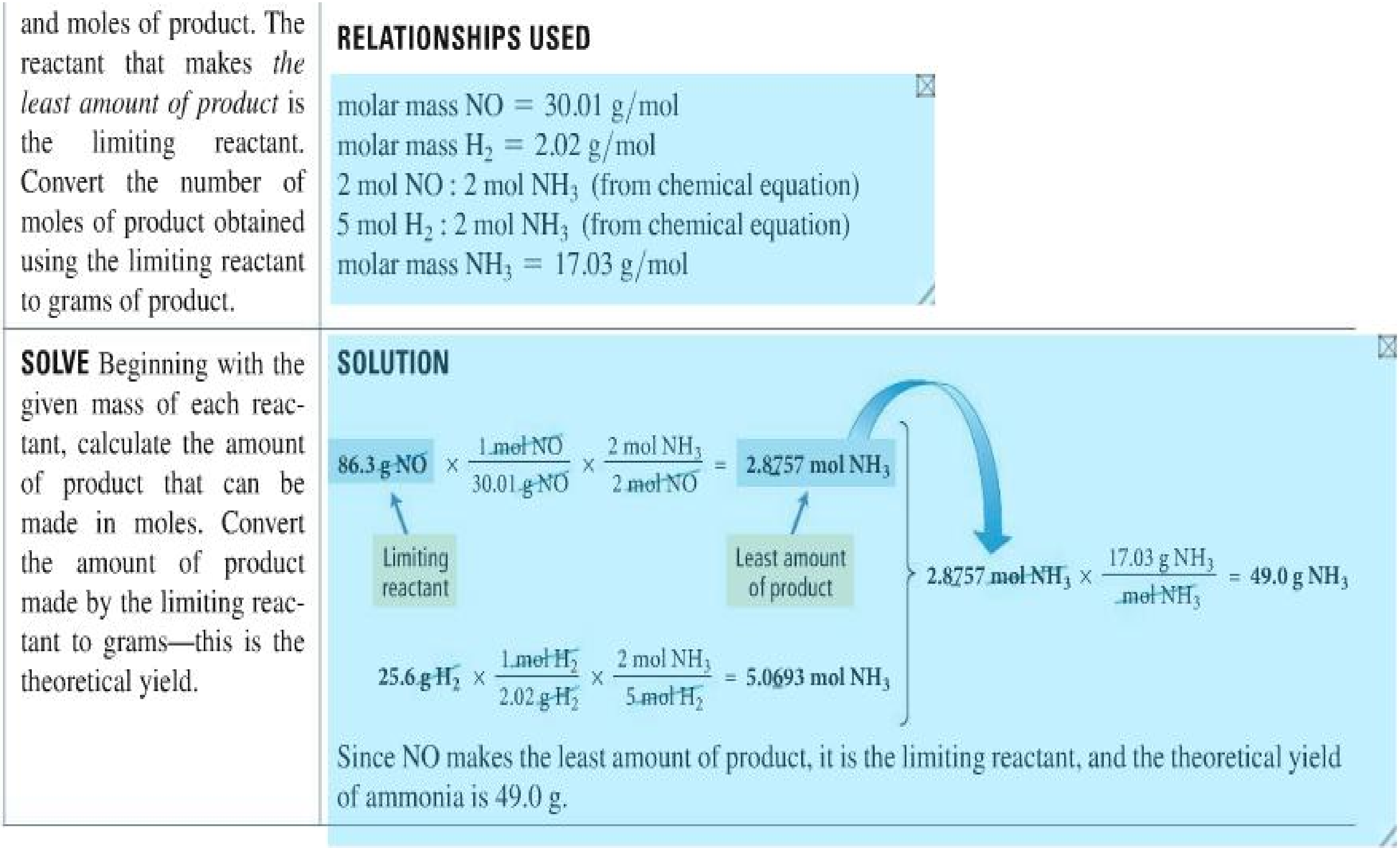


$$g\,NO \xrightarrow{\frac{1\ mol\ NO}{30.01\ g\ NO}} mol\,NO \xrightarrow{\frac{2\ mol\ NH_3}{2\ mol\ NO}} mol\,NH_3$$

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

Smallest amount determines limiting reactant.

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



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

$$86.3\,\cancel{\text{g NO}} \times \frac{1\,\cancel{\text{mol NO}}}{30.01\,\cancel{\text{g NO}}} \times \frac{2\,\text{mol NH}_3}{2\,\cancel{\text{mol NO}}} = 2.8757\,\text{mol NH}_3$$

Limiting reactant

Least amount of product

$$25.6\,\cancel{\text{g H}_2} \times \frac{1\,\cancel{\text{mol H}_2}}{2.02\,\cancel{\text{g H}_2}} \times \frac{2\,\text{mol NH}_3}{5\,\cancel{\text{mol H}_2}} = 5.0693\,\text{mol NH}_3$$

$$2.8757\,\cancel{\text{mol NH}_3} \times \frac{17.03\,\text{g NH}_3}{\cancel{\text{mol NH}_3}} = 49.0\,\text{g NH}_3$$

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

Limiting Reactant and Percent Yield: Gram to Gram

Example: **$2Na(s) + Cl_2(g) \rightarrow 2\ NaCl(s)$**

Given (***grams***): **53.2 g Na and 65.8 g Cl_2**

Find: limiting reactant and theoretical yield

SOLUTION:

$$53.2\ \cancel{\text{g Na}} \times \frac{1\ \cancel{\text{mol Na}}}{22.99\ \cancel{\text{g Na}}} \times \frac{2\ \cancel{\text{mol NaCl}}}{2\ \cancel{\text{mol Na}}} \times \frac{58.44\ \text{g NaCl}}{1\ \cancel{\text{mol NaCl}}} = 135\ \text{g NaCl}$$

$$65.8\ \cancel{\text{g Cl}_2} \times \frac{1\ \cancel{\text{mol Cl}_2}}{70.90\ \cancel{\text{g Cl}_2}} \times \frac{2\ \cancel{\text{mol NaCl}}}{1\ \cancel{\text{mol Cl}_2}} \times \frac{58.44\ \text{g NaCl}}{1\ \cancel{\text{mol NaCl}}} = \boxed{108\ \text{g NaCl}}$$

Smallest amount determines limiting reactant.

Limiting reactant

Least amount of product

Theoretical Yield and Percent Yield

Example: **$2Na(s) + Cl_2(g) \rightarrow 2\,NaCl(s)$**

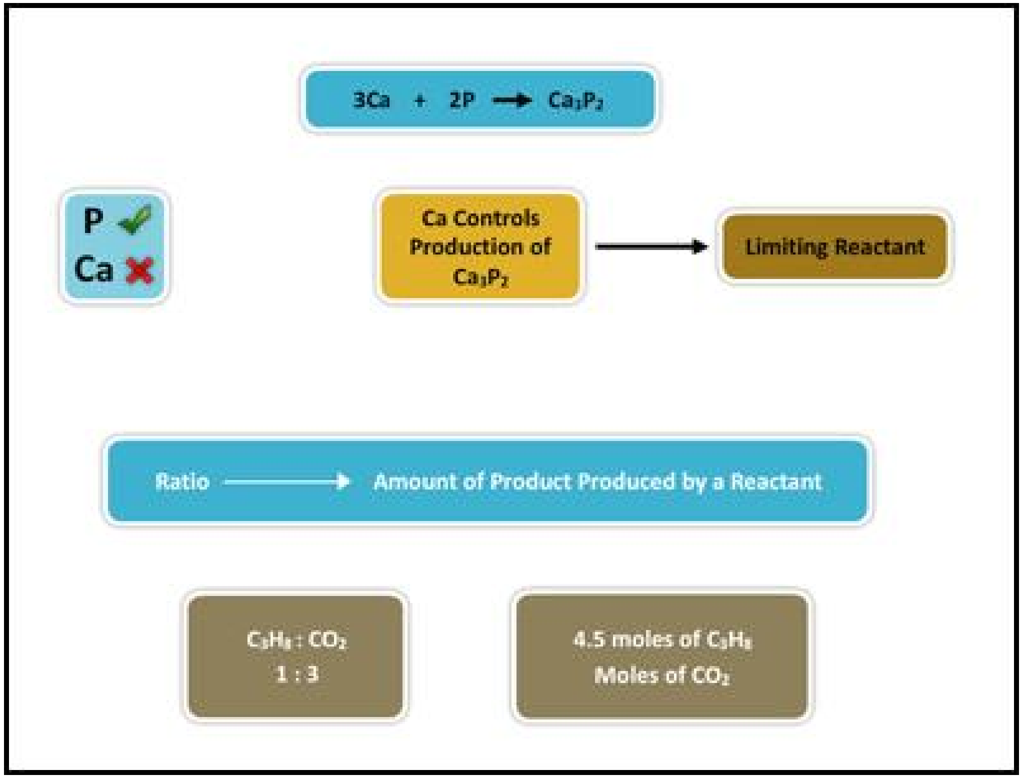
Given (***grams***): **actual yield 86.4 g NaCl**

Find: percent yield

- The actual yield is usually less than the theoretical yield because at least a small amount of product is lost or does not form during a reaction.

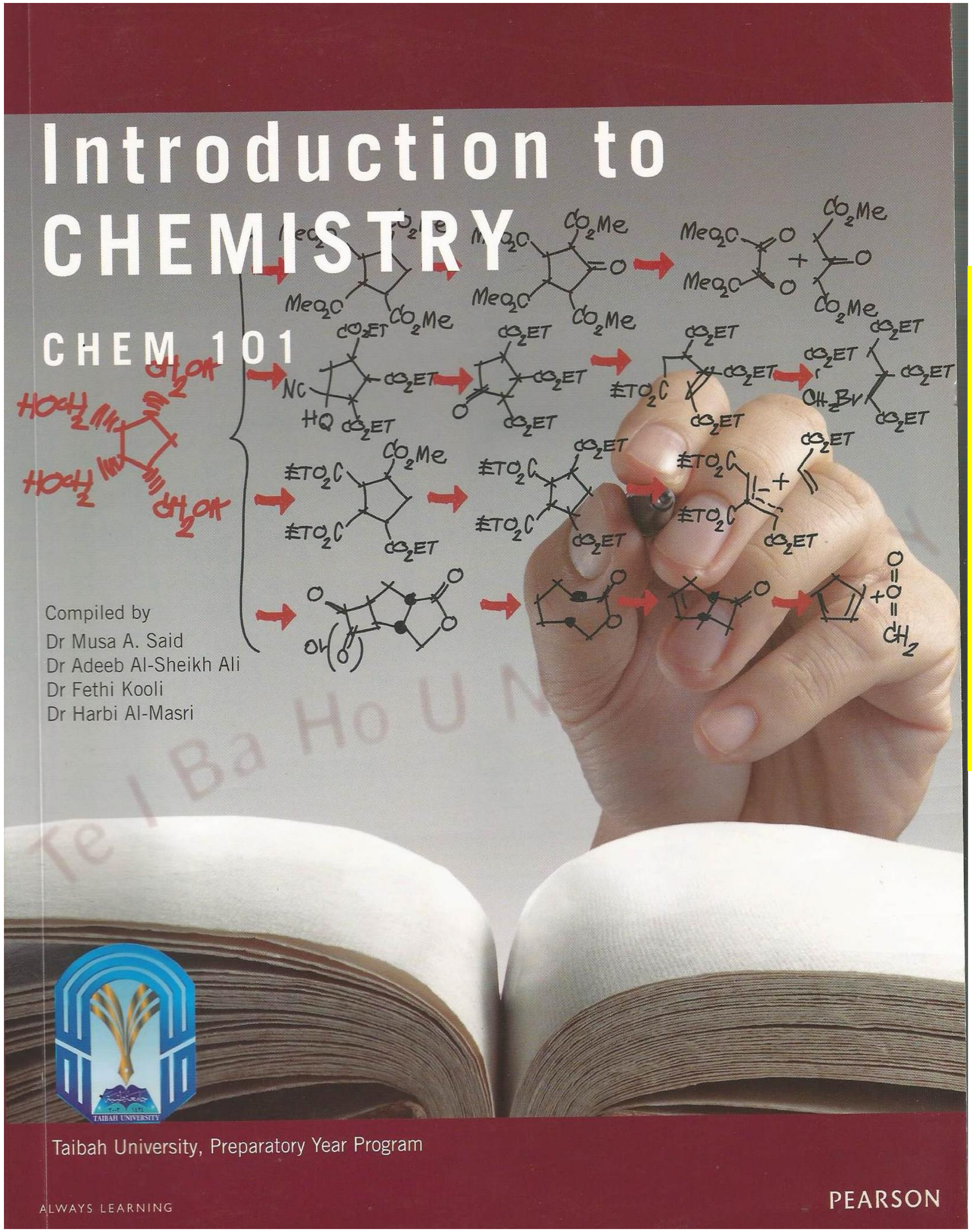
$$\boxed{\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% = \frac{86.4\ \text{g}}{108\ \text{g}} \times 100\% = 80.0\%}$$

Review



In this les-son, we have learned how to:

- Iden-tify sto-i-chio-met-ric quan-ti-ties.
- Cal-cu-late lim-it-ing re-ac-tants and the-o-ret-i-cal and per-cent yield.



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Chapter 4

CHEMICAL BONDING
CHEMICAL REACTION

TOPIC 13
SOLUTION CONCENTRATION
TYPES OF AQUEOUS SOLUTION

Introduction

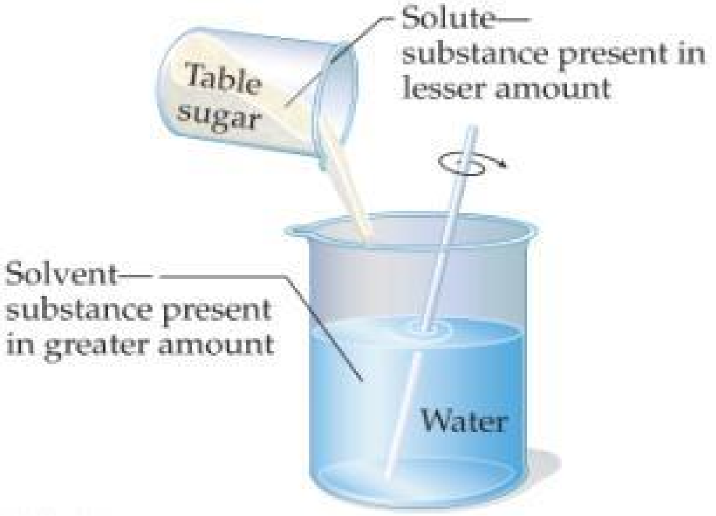
OBJECTIVES

- 4.2.1 Calculate the molarity
- 4.2.2 Deal with solution dilution
- 4.2.3 Use solution stoichiometry to find volumes and amounts

Since most chem-i-cal re-ac-tions occur in so-lu-tions, it is im-por-tant to un-der-stand how to pre-pare the proper con-cen-tra-tions of the so-lu-tions and how they af-fect the de-sired amount of each prod-uct.

What Is a Solution?

- **Solution**
  - Homogenous mixture of two or more substances
    - **Solvent**—material present in largest amount
    - **Solute**—all other materials present
  - Consider sugar dissolved in water.
    - Water is the solvent.
    - Sugar is the solute.



Solution Concentration

The amount of solute in a solution is given by its concentration.

- **Molarity shows the relationship between the moles of solute and liters of solution.**

- **It is used because it describes how many molecules of solute are in each liter of solution.**

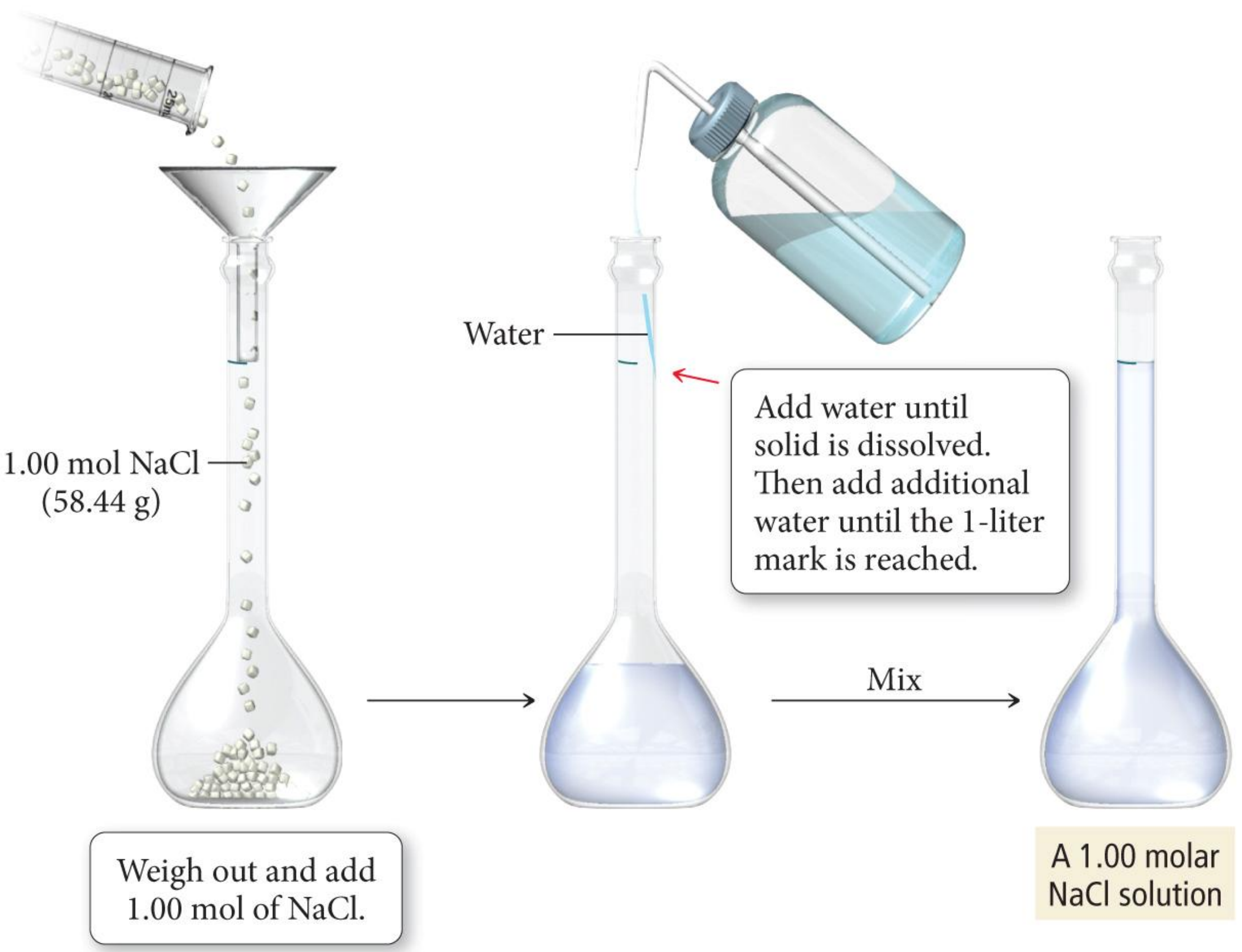
$$\text{Molarity(M)} = \frac{\text{amount of solute in moles}}{\text{amount of solution in liters}}$$

- **Concentration unit = moles of solute per liter of solution**

$$M = mol/L$$

Preparing 1 L of a 1.00 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.

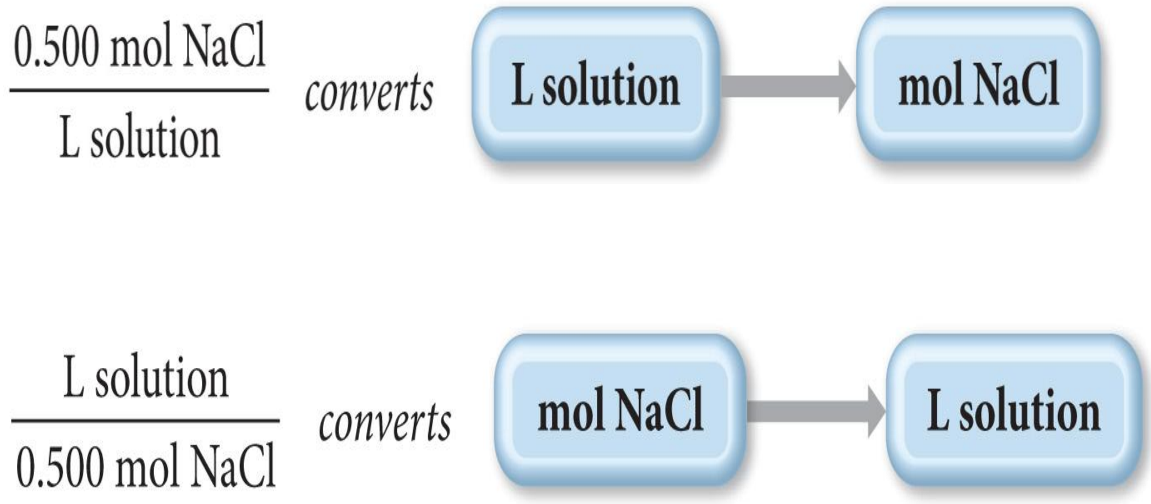
A 1.00 molar NaCl solution

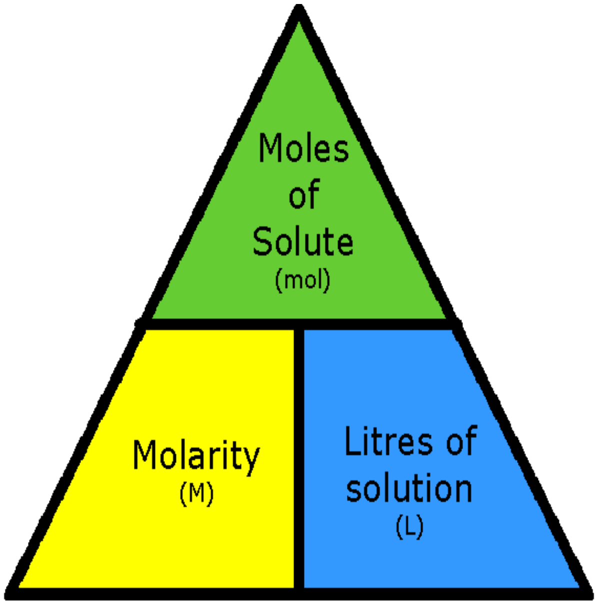
Example 4.3: Find the molarity of a solution that has 25.5 g KBr dissolved in 1.75 L of solution

Given: **Find:**	25.5 g KBr, 1.75 L solution molarity, M
Conceptual Plan: **Relationships:**	$\text{g KBr} \xrightarrow[\frac{1\,\text{mol}}{119.00\,\text{g}}]{} \text{mol KBr}$, $\text{L sol'n} \xrightarrow[\text{M} = \frac{\text{mol}}{\text{L}}]{} \text{M}$ 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

Using Molarity in Calculations

- We can use the molarity of a solution as a conversion factor between moles of the solute and liters of the solution.
  - ✓ For example, a 0.500 M NaCl solution contains 0.500 mol NaCl for every liter of solution.



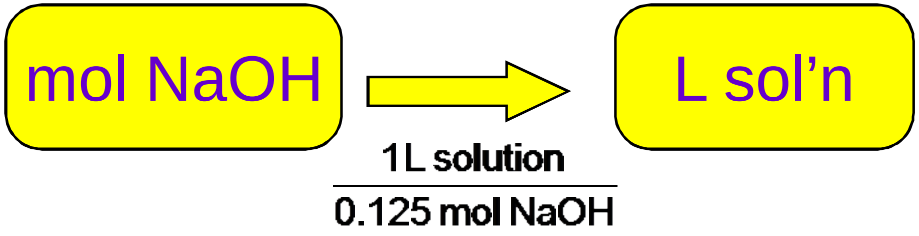


Example 4.6: How many liters of 0.125 M NaOH contain 0.255 mol NaOH?

Given: 0.125 M NaOH, 0.255 mol NaOH

Find: liters, L

Conceptual Plan:



mol NaOH → L sol'n

$$\frac{1\text{ L solution}}{0.125\text{ mol NaOH}}$$

Relationships: 0.125 mol NaOH = 1 L solution

Solution:

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

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

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:

    moles solute in solution 1 = moles solute in solution 2
- The concentrations and volumes of the stock and new solutions are inversely proportional:

$$M_1 \cdot V_1 = M_2 \cdot V_2$$

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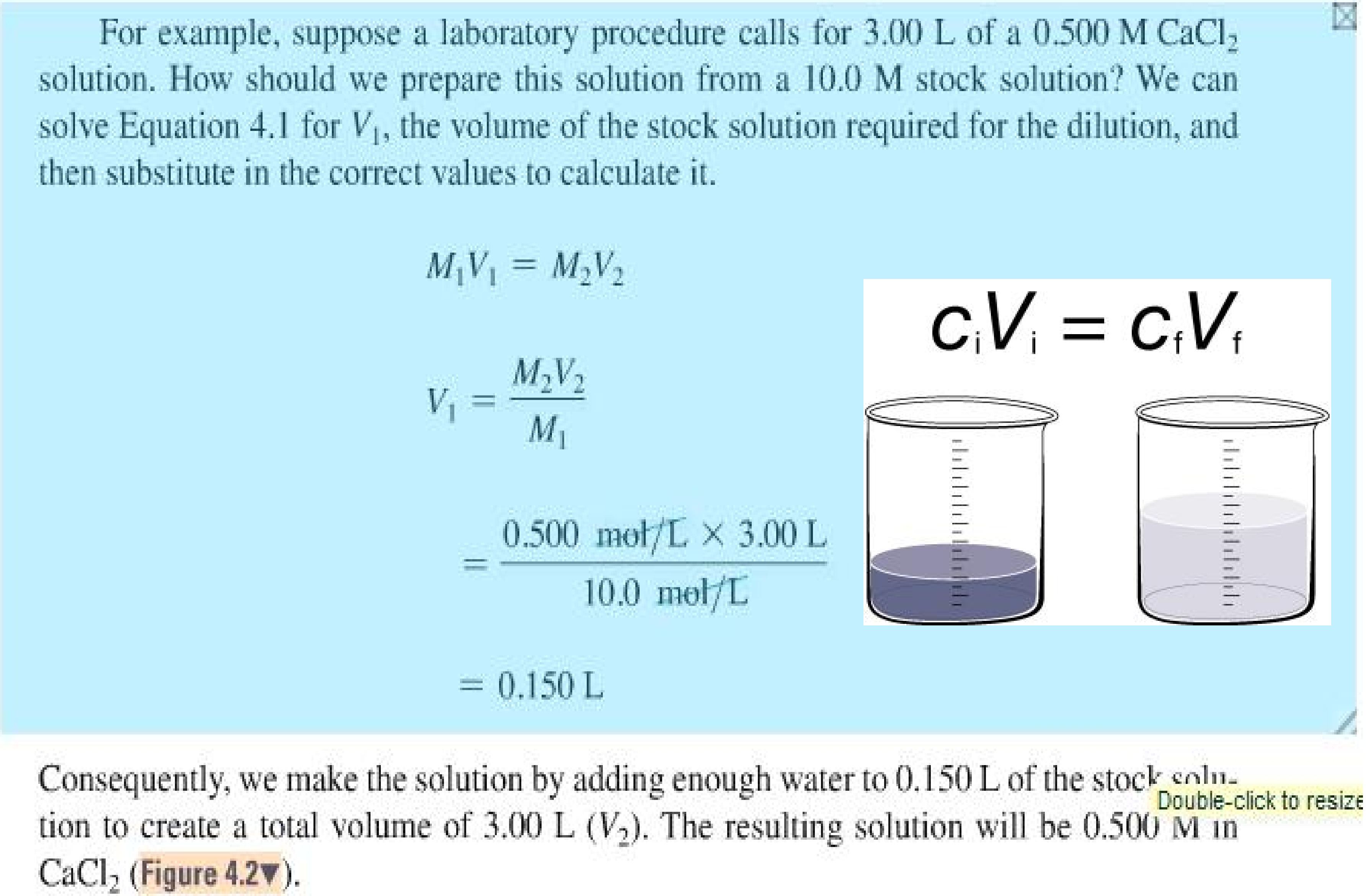


For example, suppose a laboratory procedure calls for 3.00 L of a 0.500 M $CaCl_2$ solution. How should we prepare this solution from a 10.0 M stock solution? We can solve Equation 4.1 for V_1, the volume of the stock solution required for the dilution, and then substitute in the correct values to calculate it.

$$M_1V_1 = M_2V_2$$

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

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

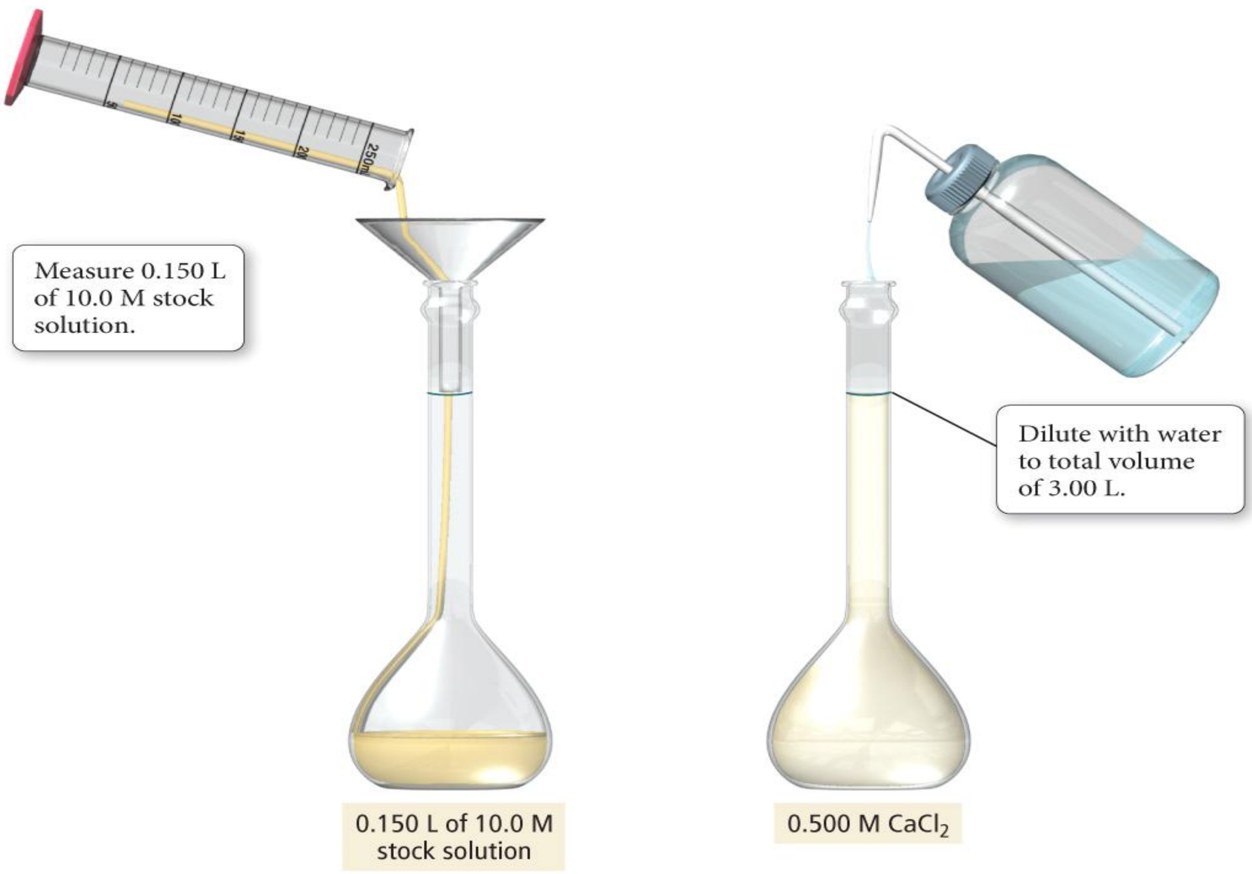
$$= 0.150\ \text{L}$$

$$c_iV_i = c_fV_f$$

Consequently, we make the solution by adding enough water to 0.150 L of the stock solution to create a total volume of 3.00 L (V_2). The resulting solution will be 0.500 M in $CaCl_2$ (**Figure 4.2▼**).

Preparing 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 $CaCl_2$

$$M_1V_1 = M_2V_2$$

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

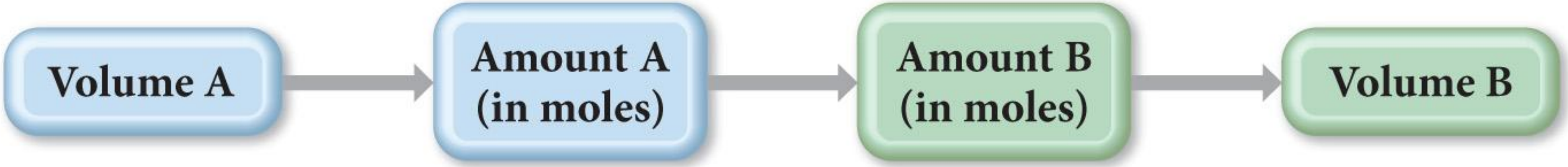
$$1.50\text{ mol} = 1.50\text{ mol}$$

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

Given: **Find:**	$M_1 = 2.0\ \text{M}, M_2 = 0.25\ \text{M}, V_2 = 200.0\ \text{mL}$ V_1, L
Conceptual Plan:	$M_1, M_2, V_2 \longrightarrow V_1$ $V_1 = \frac{M_2 \bullet V_2}{M_1}$
Relationships:	$M_1V_1 = M_2V_2$
Solution:	$\frac{\left(0.25\ \frac{\cancel{\text{mol}}}{\text{L}}\right) \bullet (200.0\ \text{mL})}{\left(2.0\ \frac{\cancel{\text{mol}}}{\text{L}}\right)} = 25\ \text{mL}$ 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

Solution Stoichiometry

- Because molarity relates the moles of solute to the liters of solution, it can be used to convert between amount of reactants and/or products in a chemical reaction.
  - ✓ The general conceptual plan for these kinds of calculations begins with the volume of a reactant or product.



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?

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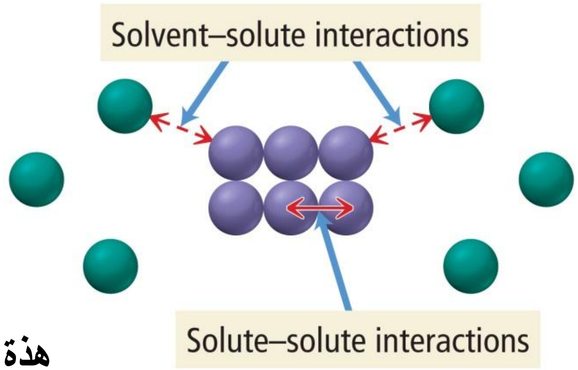
What Happens When a Solute Dissolves?

- 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.

Solute and Solvent Interactions



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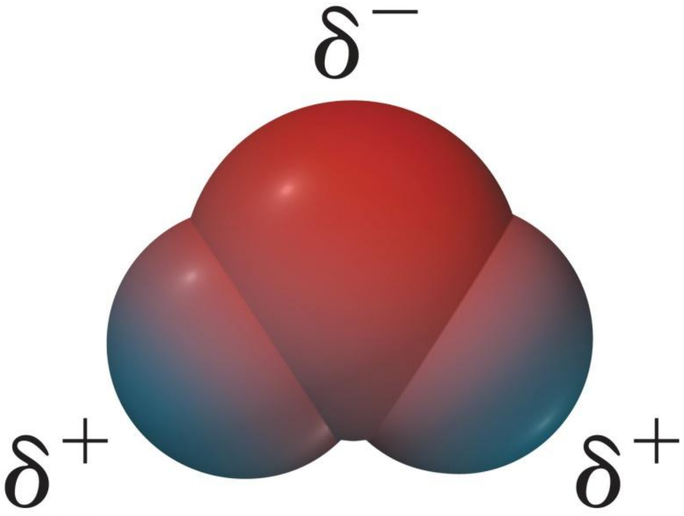
Charge Distribution in a Water Molecule

- 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^+).



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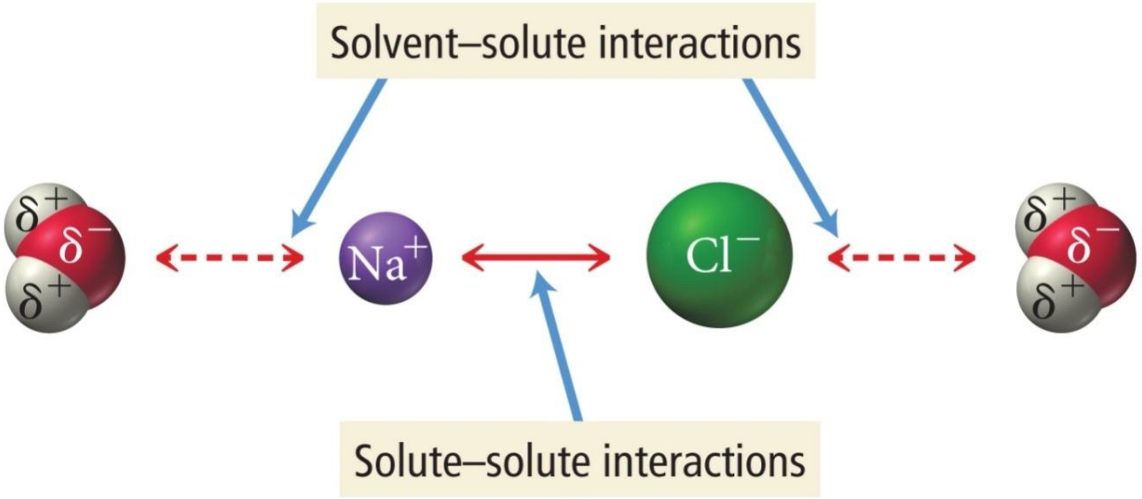
Solute and Solvent Interactions in a Sodium Chloride Solution

- 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.

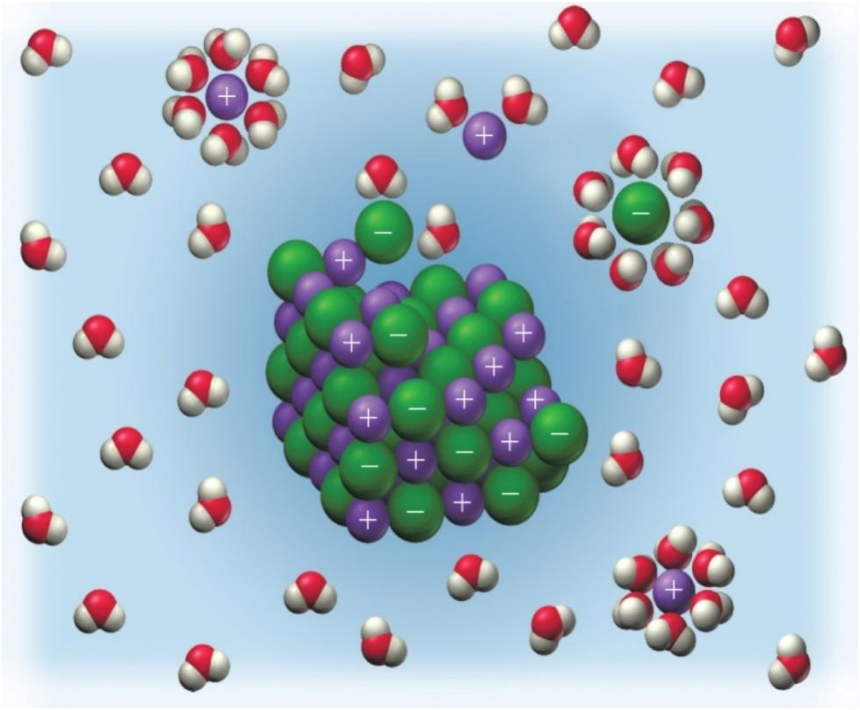


Interactions in a Sodium Chloride Solution

Sodium Chloride Dissolving in Water

- Each ion is attracted to the surrounding water molecules and pulled off and away from the crystal.
- When it enters the solution, the ion is surrounded by water molecules, insulating it from other ions.
- The result is a solution with free moving charged particles able to conduct electricity.

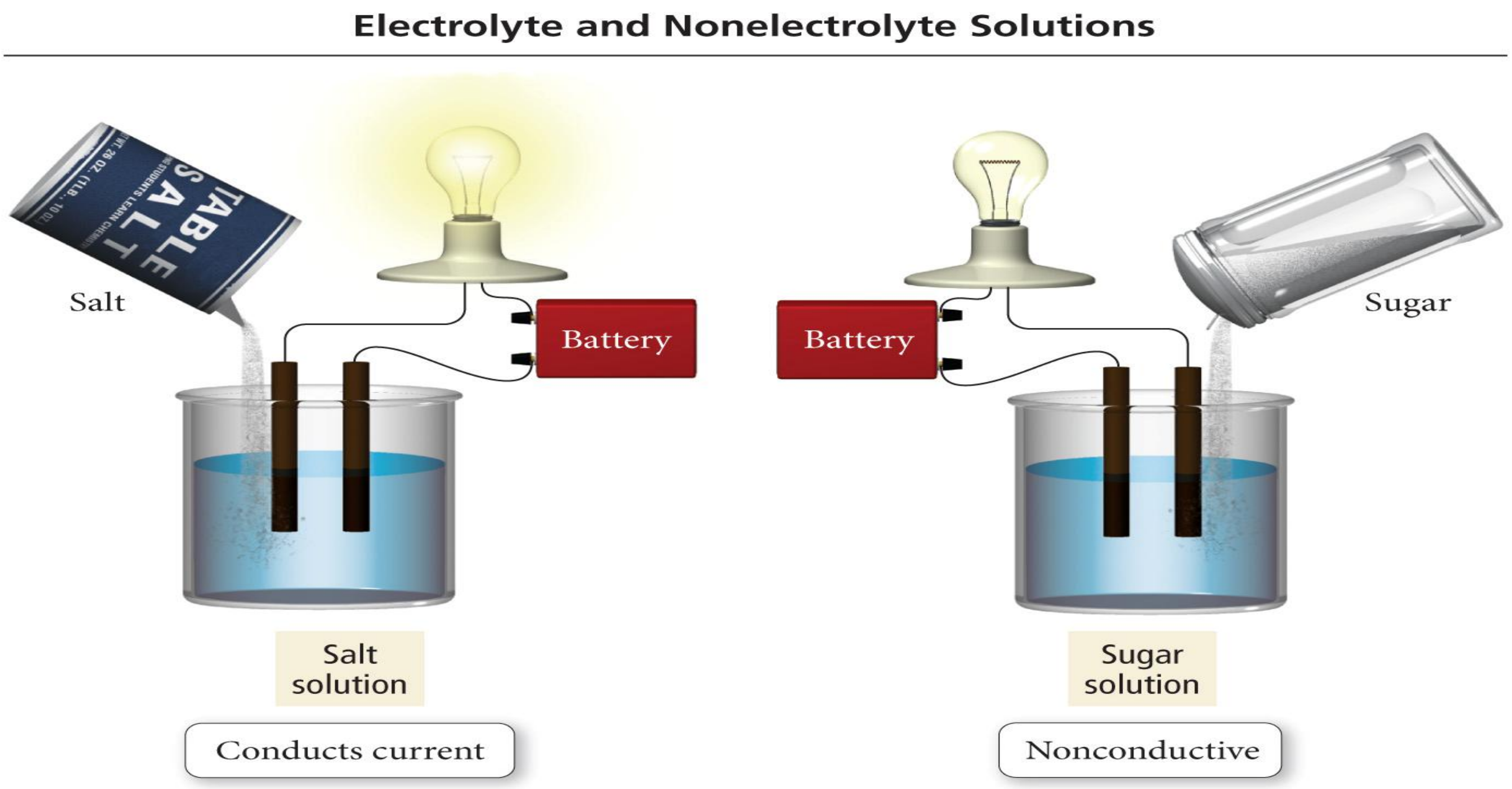
Dissolution of an Ionic Compound



Electrolyte and Nonelectrolyte Solutions

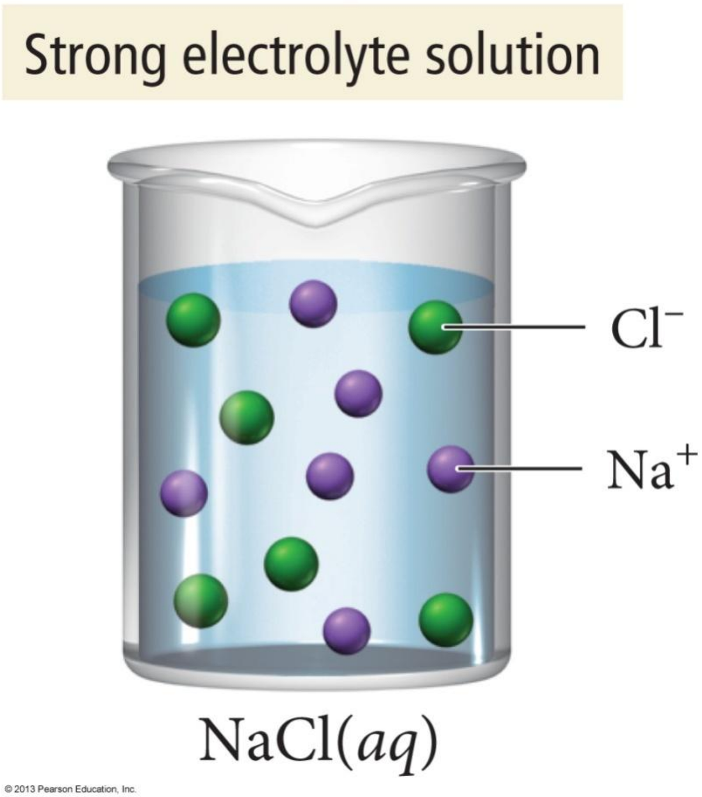
- **Electrolytes can be described as:**
  - **Strong Electrolyte:**
    - **Chemical substances that IONIZE COMPLETELY into their ions**
      - **Examples: Soluble salts and strong acids or bases**
      - $HCl\ (aq) \rightarrow H^+ + Cl^-$ **or** $CuCl_2 \rightarrow Cu^{2+} + 2\ Cl^-$
    - **Can conduct electrical current**
  - **Weak Electrolyte:**
    - **Chemical substances that IONIZE PARTIALLY into their ions**
      - **Examples: weak acids or weak bases**
        - $CH_3COOH(aq) \rightarrow CH_3COO^-\ (aq) + H^+(aq)$
    - **Can conduct electrical current**
  - **Nonelectrolytes:**
    - **Chemical substances that DISSOLVE in water but NOT AS IONS**
    - **They do not conduct electricity.**
      - **Example: polar substances such as sugar or alcohol**
        - $C_6H_{12}O_6(s) \rightarrow C_6H_{12}O_6(aq)$

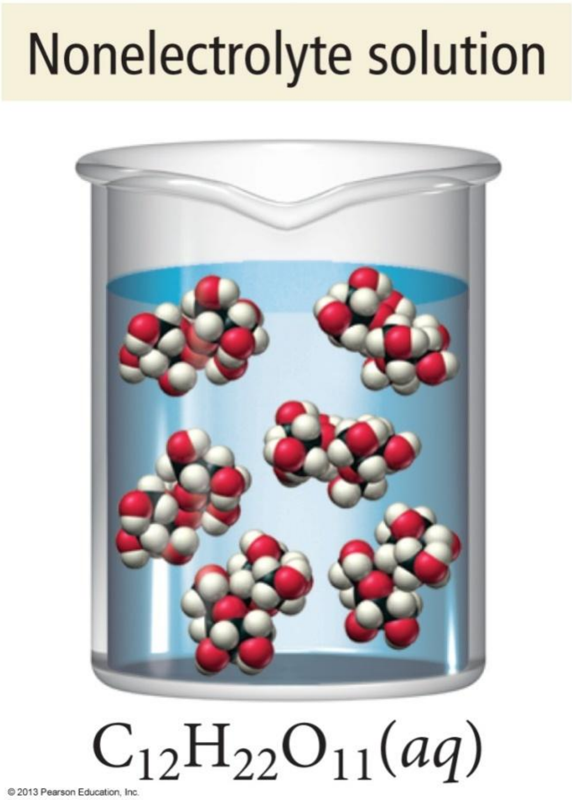
Electrolytes and Nonelectrolytes



- **A solution of salt (an electrolyte) conducts electrical currentnot.**
- **A solution of sugar (a nonelectrolyte) does not.**

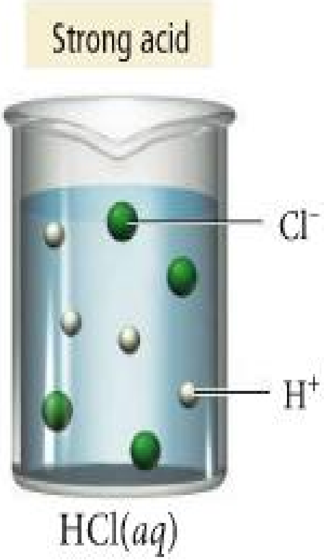
Salt (Ionic compound) vs. Sugar Dissolved in Water



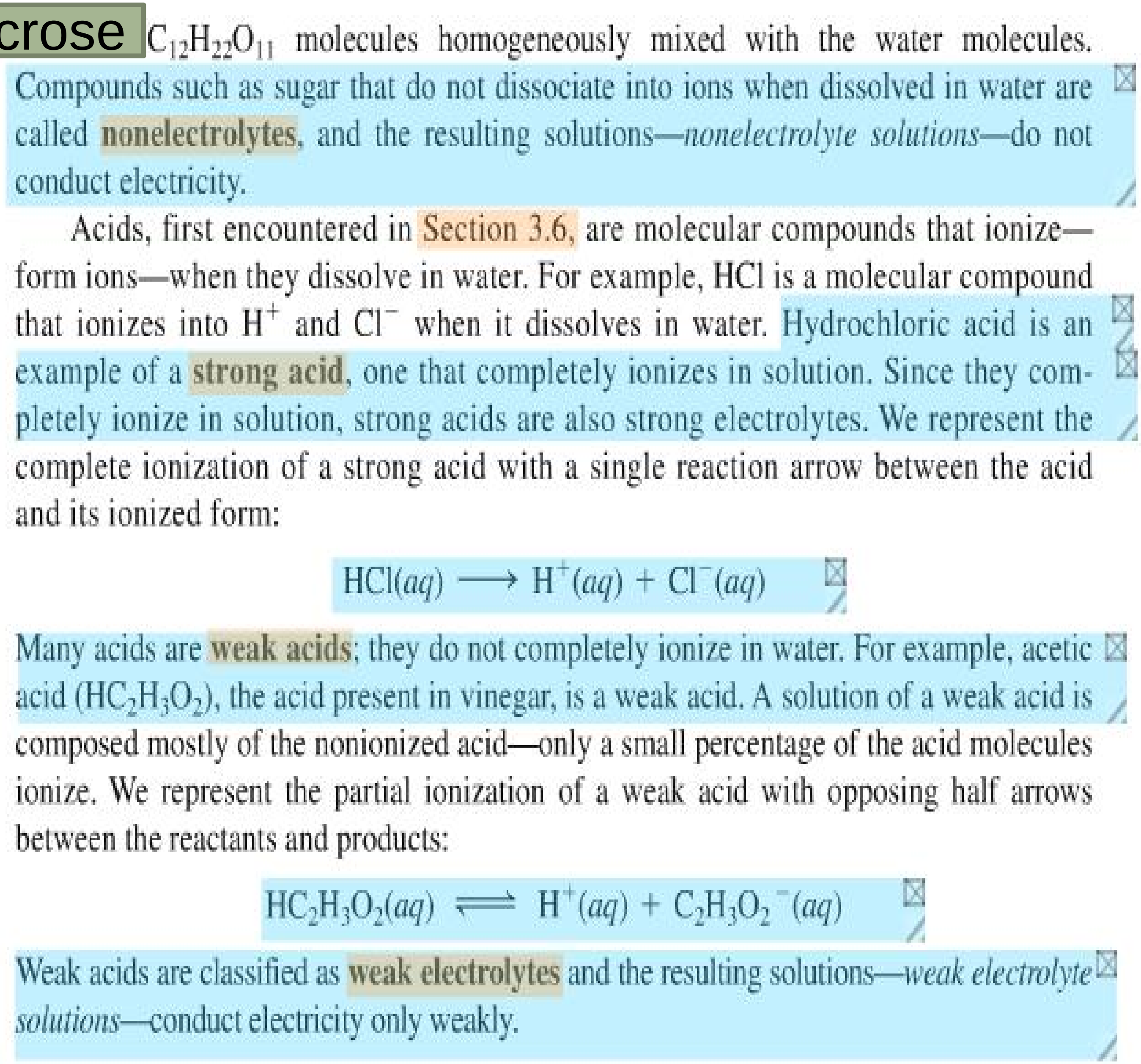


Salts (ionic compound) ionize when dissolved in water to form ions

Molecular compounds do not dissociate when they dissolve; they DO NOT form ions.



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—*nonelectrolyte solutions*—do not conduct electricity.

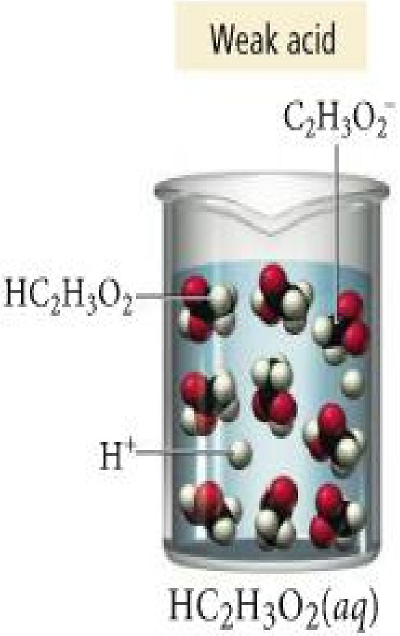
Acids, first encountered in Section 3.6, are molecular compounds that ionize—form ions—when they dissolve in water. For example, HCl is a molecular compound that ionizes into H^+ and Cl^- when it dissolves in water. Hydrochloric acid is an example of a **strong acid**, one that completely ionizes in solution. Since they completely ionize in solution, strong acids are also strong electrolytes. We represent the complete ionization of a strong acid with a single reaction arrow between the acid and its ionized form:

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

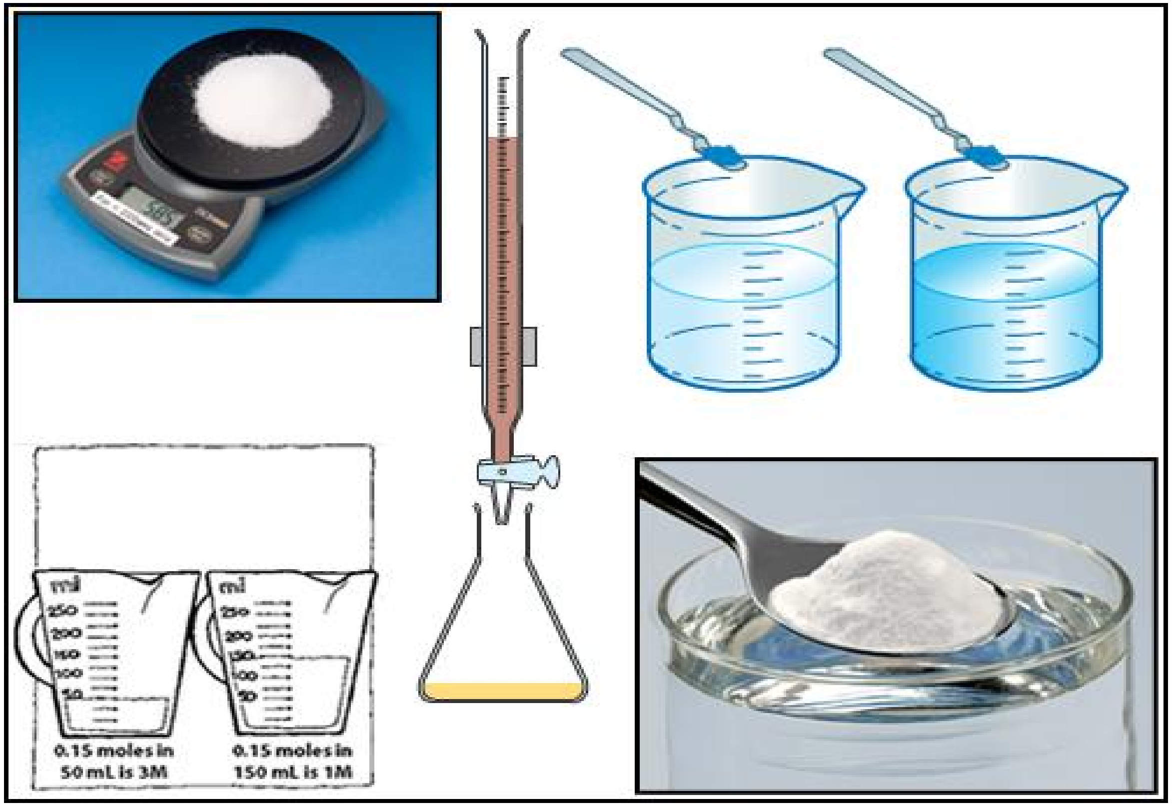
Many acids are **weak acids**; they do not completely ionize in water. For example, acetic acid ($HC_2H_3O_2$), the acid present in vinegar, is a weak acid. A solution of a weak acid is composed mostly of the nonionized acid—only a small percentage of the acid molecules ionize. We represent the partial ionization of a weak acid with opposing half arrows between the reactants and products:

$$HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

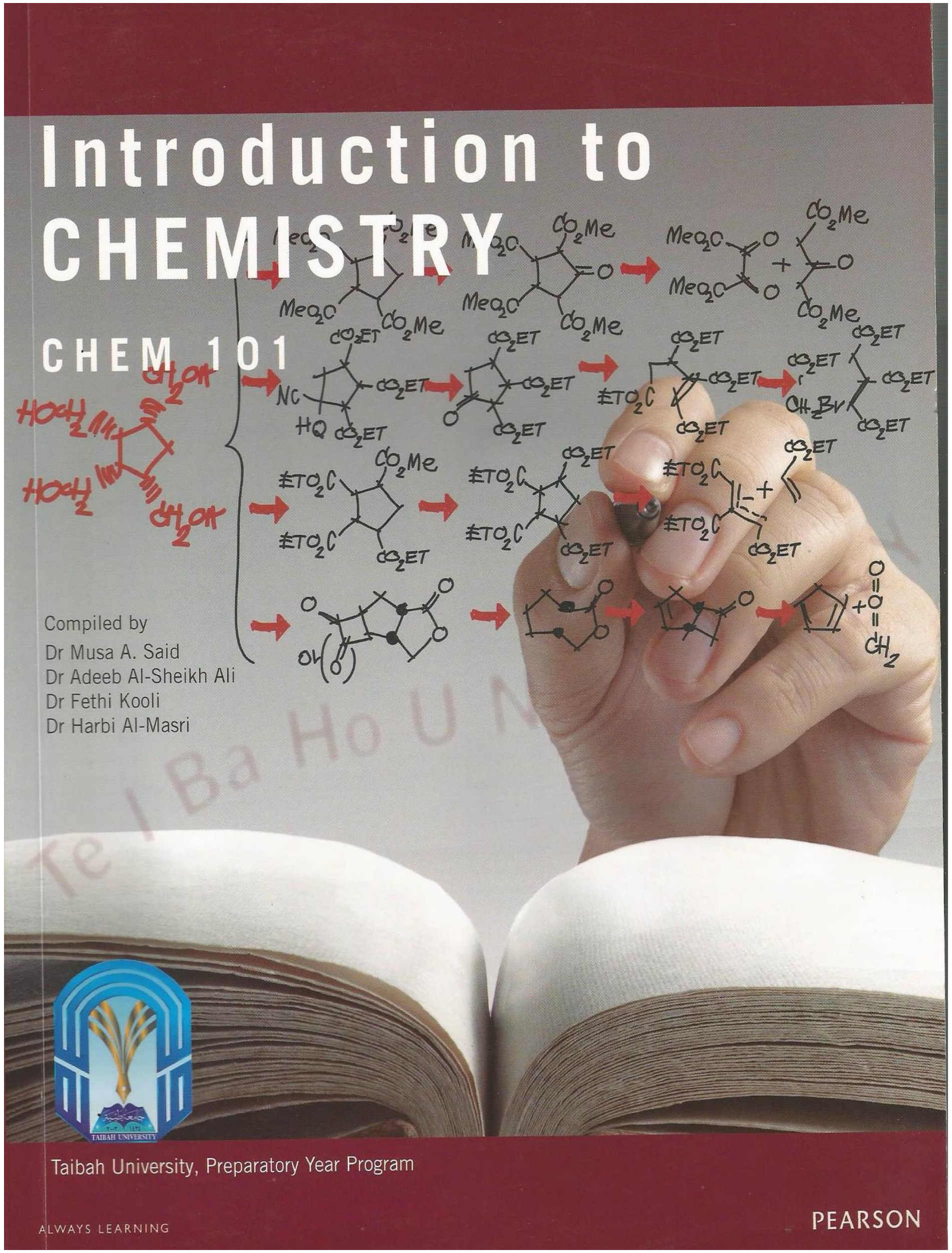
Weak acids are classified as **weak electrolytes** and the resulting solutions—*weak electrolyte solutions*—conduct electricity only weakly.



Review



In this les-son, we have learned how to:
- Cal-cu-late the mo-lar-ity
- Deal with so-lu-tion di-lu-tion
- Use so-lu-tion sto-i-chiom-e-try to find vol-umes and amounts



Introduction to
CHEMISTRY
CHEM 101

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ALWAYS LEARNING

PEARSON

Lecture Presentation

Chapter 4

CHEMICAL BONDING
CHEMICAL REACTION

TOPIC 14

Acid−Base Reactions
Oxidation Reduction Reaction

Introduction

Acid-Base Reactions

Reactions between acids and bases are very common in chemistry and in everyday life, as are oxidation reactions in which the oxidation number of an element changes.

OBJECTIVES

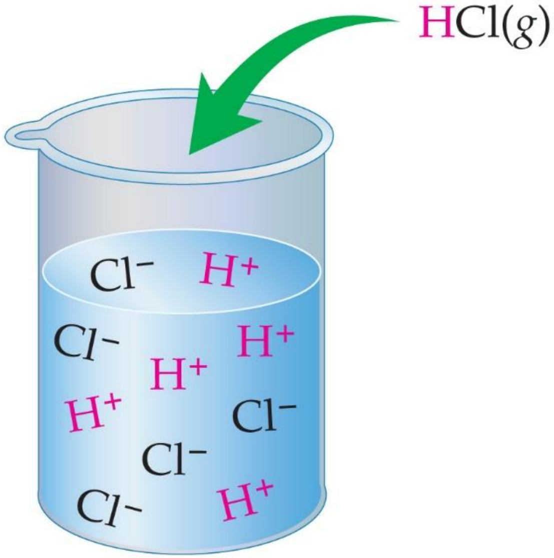
4.1.1 Writing equations for acid-base reactions

4.1.2 Calculate oxidation state

4.1.3 Identify oxidant and reductant using oxidation state

4.7 Acid–Base Reactions

- **Acid**
  - ✓ Any compound that dissociates to produce H^+ in water
  - ✓ Examples include acetic acid, HCl, and HF.



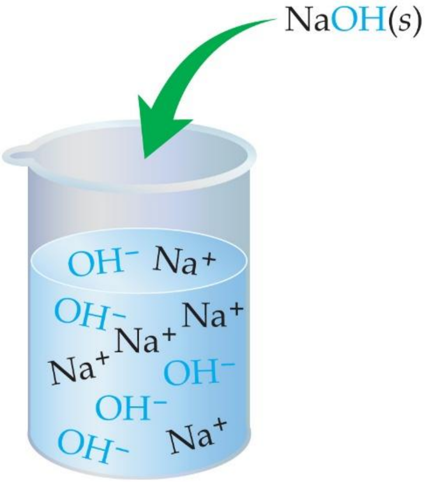
HCl is an acid because it dissociates in water to produce $H^+(aq)$ ions.

MP4

anim_Ionically Bonded Acids_22_02.mp4

4.7 Acid–Base Reactions

- **Base**
  - ✓ Substance that dissociates to produce OH^- in water
  - ✓ Examples include NaOH, KOH, and $Ca(OH)_2$.



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

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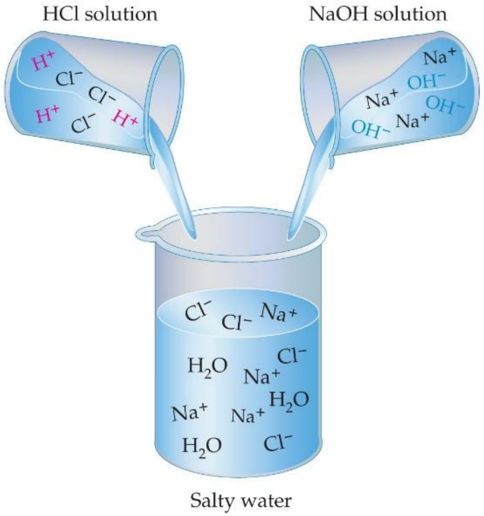
4.7 Acid–Base Reactions

- **Acid–base neutralization**

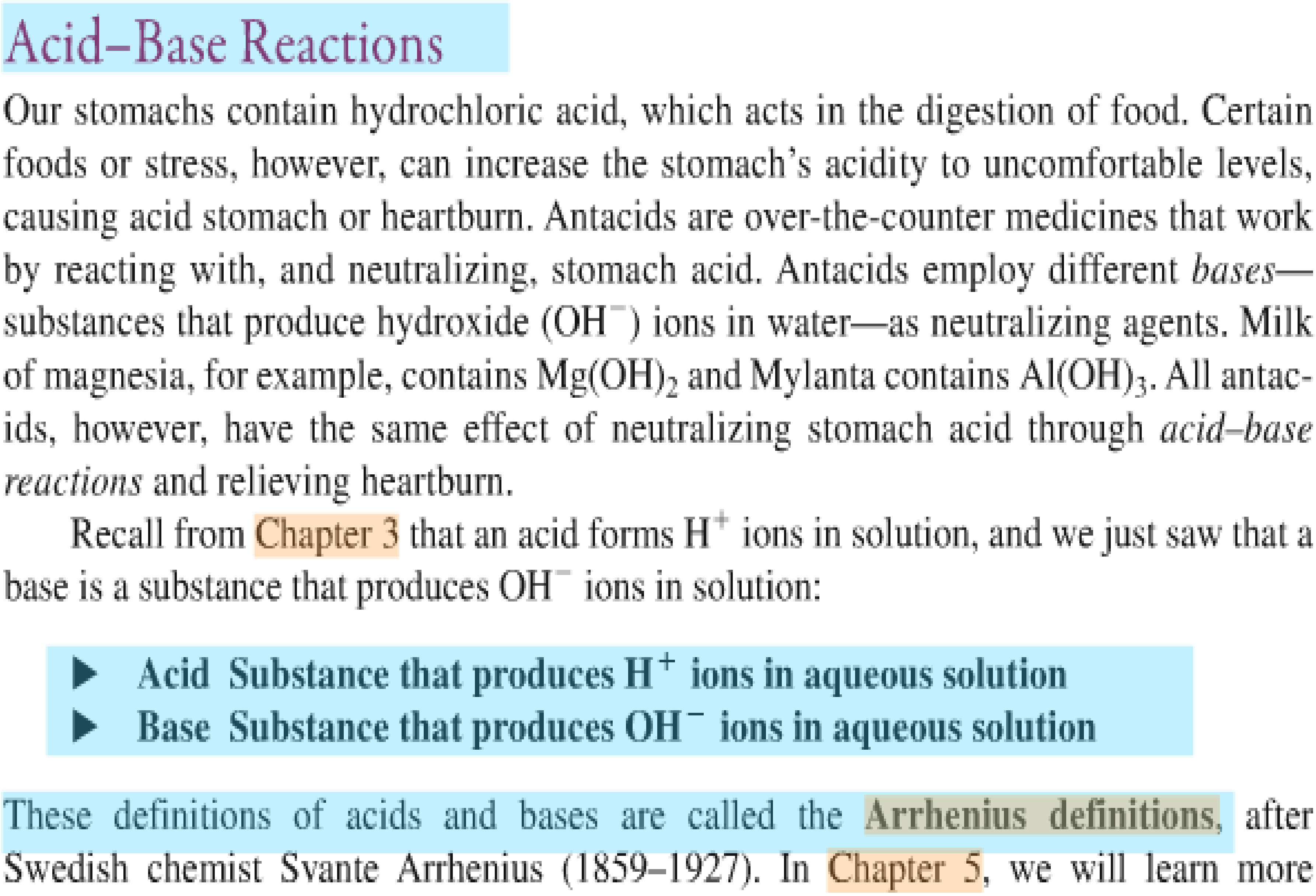


anim_Neutralization Reactions_22_01.mp4

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

- Net ionic equation

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



Acid–Base Reactions

Our stomachs contain hydrochloric acid, which acts in the digestion of food. Certain foods or stress, however, can increase the stomach's acidity to uncomfortable levels, causing acid stomach or heartburn. Antacids are over-the-counter medicines that work by reacting with, and neutralizing, stomach acid. Antacids employ different *bases*—substances that produce hydroxide (OH^-) ions in water—as neutralizing agents. Milk of magnesia, for example, contains $Mg(OH)_2$ and Mylanta contains $Al(OH)_3$. All antacids, however, have the same effect of neutralizing stomach acid through *acid–base reactions* and relieving heartburn.

Recall from Chapter 3 that an acid forms H^+ ions in solution, and we just saw that a base is a substance that produces OH^- ions in solution:

- **Acid** **Substance that produces H^+ ions in aqueous solution**
- **Base** **Substance that produces OH^- ions in aqueous solution**

These definitions of acids and bases are called the **Arrhenius definitions**, after Swedish chemist Svante Arrhenius (1859–1927). In Chapter 5, we will learn more

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According to the Arrhenius definition, HCl is an acid because it produces H^+ ions in solution:

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

An H^+ ion is a bare proton. Protons associate with water molecules in solution to form **hydronium ions** (**Figure 4.10**▶):

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

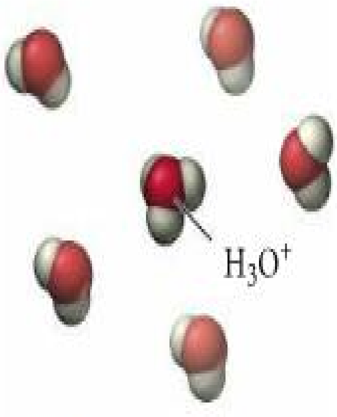
Chemists use $H^+(aq)$ and $H_3O^+(aq)$ interchangeably to mean the same thing—an H^+ ion dissolved in water. The ionization of HCl and other acids is often written to show the association of the proton with a water molecule to form the hydronium ion:

$$HCl(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

Some acids—called **polyprotic acids**—contain more than one ionizable proton and release them sequentially. For example, sulfuric acid, H_2SO_4, is a **diprotic acid**. It is strong in its first ionizable proton, but weak in its second:

$$H_2SO_4(aq) \longrightarrow H^+(aq) + HSO_4^-(aq)$$

$$HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$$



▲ **FIGURE 4.10 The Hydronium Ion**
Protons normally associate with water molecules in solution to form H_3O^+ ions, which in turn interact with other water molecules.

▲ Acids are found in lemons, limes, and vinegar. Vitamin C and aspirin are also acids.

According to the the Arrhenius definition, NaOH is a base because it produces OH^- ions in solution:

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

In analogy to diprotic acids, some bases, such as $Sr(OH)_2$, for example, produce two moles of OH^- per mole of the base.

$$Sr(OH)_2(aq) \longrightarrow Sr^{2+}(aq) + 2\,OH^-(aq)$$

Common acids and bases are listed in Table 4.1. Acids and bases are in many everyday substances.

TABLE 4.1 Some Common Acids and Bases

Name of Acid	Formula	Name of Base	Formula
Hydrochloric acid	HCl	Sodium hydroxide	$NaOH$
Hydrobromic acid	HBr	Lithium hydroxide	$LiOH$
Hydroiodic acid	HI	Potassium hydroxide	KOH
Nitric acid	HNO_3	Calcium hydroxide	$Ca(OH)_2$
Sulfuric acid	H_2SO_4	Barium hydroxide	$Ba(OH)_2$
Perchloric acid	$HClO_4$	Ammonia*	NH_3 (weak base)
Acetic acid	$HC_2H_3O_2$ (weak acid)		
Hydrofluoric acid	HF (weak acid)		

*Ammonia does not contain OH^-, but it produces OH^- in a reaction with water that occurs only to a small extent: $NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$.

▲ Many common household products contain bases.

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

Acid–base reactions generally form water and an ionic compound—called a **salt**—that usually remains dissolved in the solution. The net ionic equation for many acid–base reactions is

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

Another example of an acid–base reaction is the reaction between sulfuric acid and potassium hydroxide:

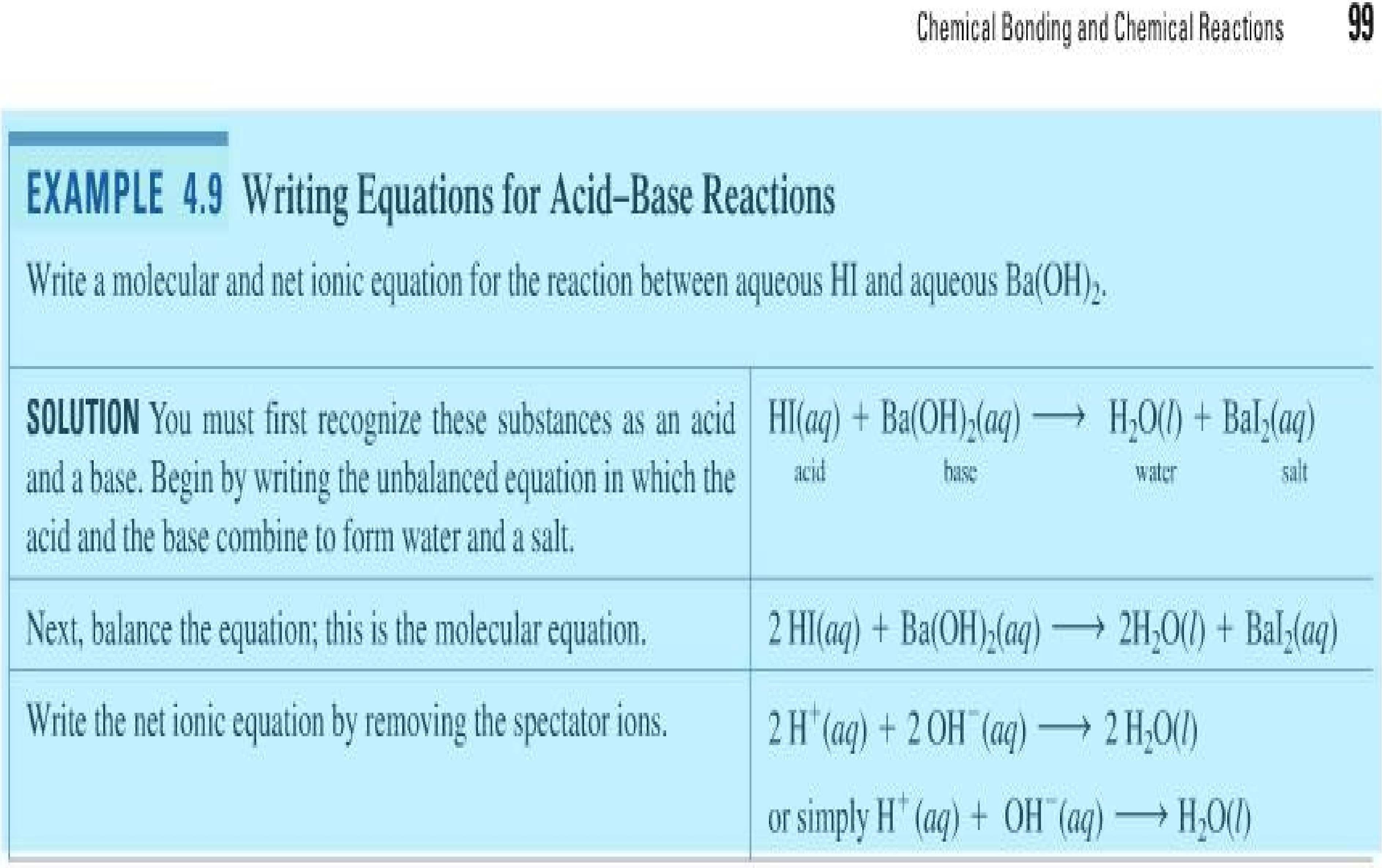
$$\underset{\text{acid}}{H_2SO_4(aq)} + \underset{\text{base}}{2\,KOH(aq)} \longrightarrow \underset{\text{water}}{2\,H_2O(l)} + \underset{\text{salt}}{K_2SO_4(aq)}$$

nim_Acid and Base Combine-22_04.mp

Again, notice the pattern of acid and base reacting to form water and a salt.

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

When writing equations for acid–base reactions, write the formula of the salt using the procedure for writing formulas of ionic compounds from Section 3.5.



EXAMPLE 4.9 Writing Equations for Acid–Base Reactions

Write a molecular and net ionic equation for the reaction between aqueous HI and aqueous $Ba(OH)_2$.

SOLUTION You must first recognize these substances as an acid and a base. Begin by writing the unbalanced equation in which the acid and the base combine to form water and a salt.	$\underset{\text{acid}}{HI(aq)} + \underset{\text{base}}{Ba(OH)_2(aq)} \longrightarrow \underset{\text{water}}{H_2O(l)} + \underset{\text{salt}}{BaI_2(aq)}$
Next, balance the equation; this is the molecular equation.	$2\,HI(aq) + Ba(OH)_2(aq) \longrightarrow 2H_2O(l) + BaI_2(aq)$
Write the net ionic equation by removing the spectator ions.	$2\,H^+(aq) + 2\,OH^-(aq) \longrightarrow 2\,H_2O(l)$ or simply $H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$

4.8 Oxidation–Reduction Reactions

Oxidation–reduction reactions or redox reactions are reactions in which electrons are transferred from one reactant to the other. The rusting of iron, the bleaching of hair, and the production of electricity in batteries involve redox reactions. Many redox reactions involve the reaction of a substance with oxygen (**Figure 4.11▼**):

$$4\,Fe(s) + 3\,O_2(g) \longrightarrow 2\,Fe_2O_3(s) \quad \text{(rusting of iron)}$$
$$2\,C_8H_{18}(l) + 25\,O_2(g) \longrightarrow 16\,CO_2(g) + 18\,H_2O(g) \quad \text{(combustion of octane)}$$
$$2\,H_2(g) + O_2(g) \longrightarrow 2\,H_2O(g) \quad \text{(combustion of hydrogen)}$$

Redox reactions need not involve oxygen. Consider, for example, the reaction between sodium and chlorine to form sodium chloride (NaCl):

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

This reaction is similar to the reaction between sodium and oxygen to form sodium oxide:

$$4\,Na(s) + O_2(g) \longrightarrow 2\,Na_2O(s)$$

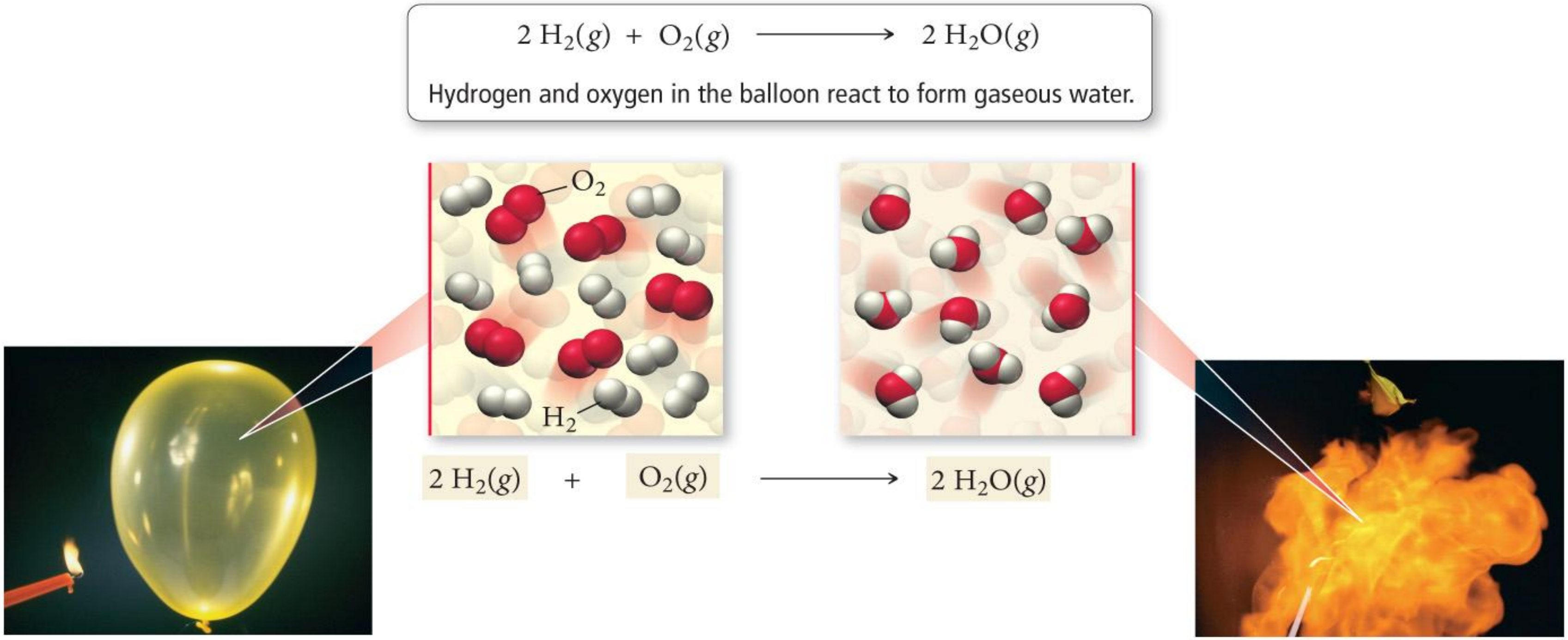
In both cases, a metal (which has a tendency to lose electrons) reacts with a nonmetal (which has a tendency to gain electrons). In both cases, metal atoms lose electrons to nonmetal atoms. A fundamental definition of **oxidation** is the loss of electrons, and a fundamental definition of **reduction** is the gain of electrons.

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.

Oxidation–Reduction Reaction



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

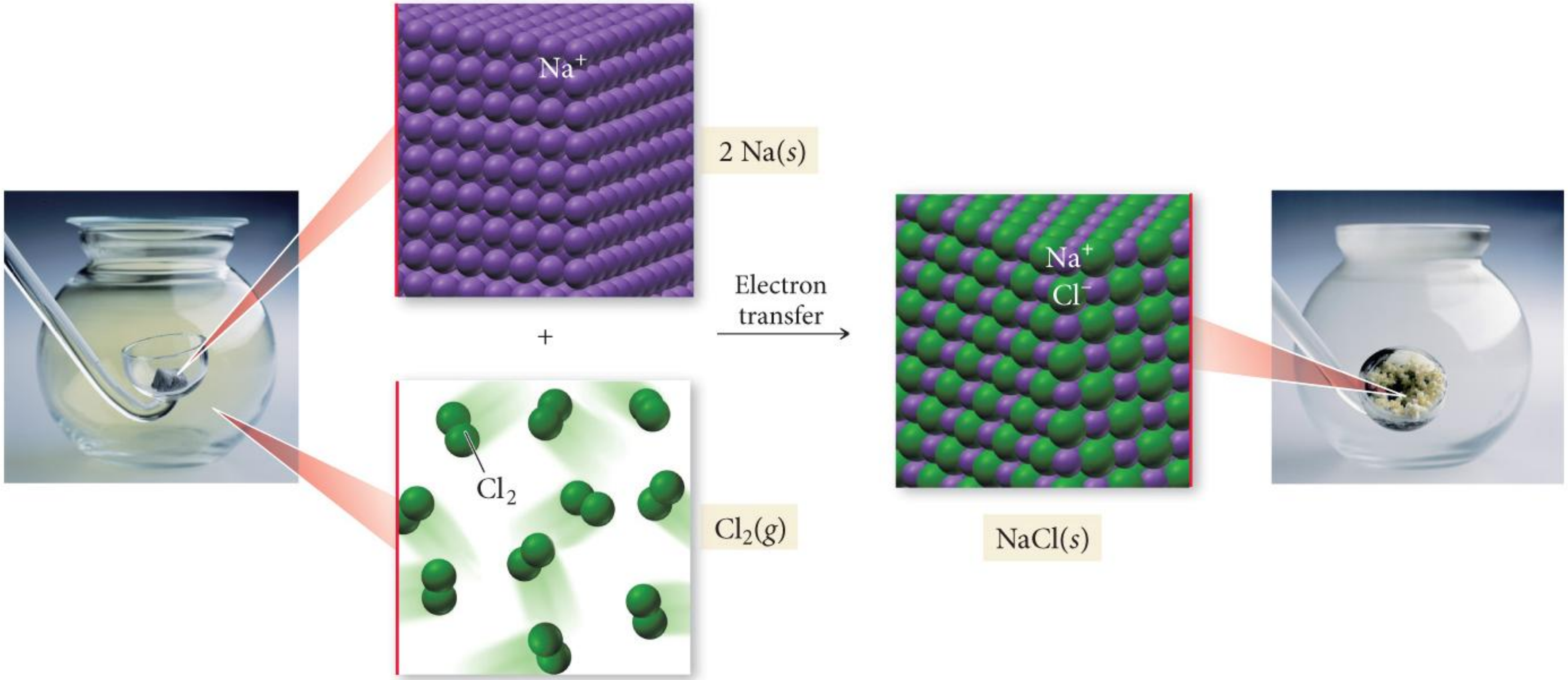
Hydrogen and oxygen in the balloon react to form gaseous water.

- Loss of electrons is **oxidation**.
- Gain of electrons is **reduction**.
- One cannot occur without the other.
- The reactions are often called **redox reactions**.

Oxidation–Reduction Reaction without Oxygen

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

Electrons are transferred from sodium to chlorine, forming sodium chloride.
Sodium is oxidized and chlorine is reduced.



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.

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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.		

Oxidation States of Nonmetals

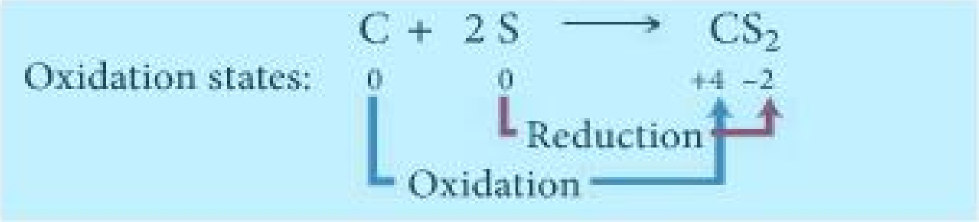
Nonmetal	Oxidation State	Example
Fluorine	–1	MgF_2 –1 ox state
Hydrogen	+1	H_2O +1 ox state
Oxygen	–2	CO_2 –2 ox state
Group 7A	–1	CCl_4 –1 ox state
Group 6A	–2	H_2S –2 ox state
Group 5A	–3	NH_3 –3 ox state

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.



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

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.

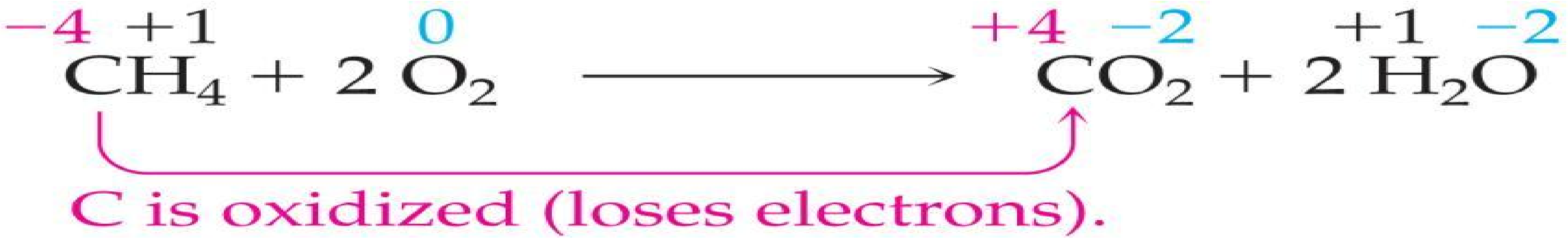
- **Oxidation** **An increase in oxidation state**
- **Reduction** **A decrease in oxidation state**

- Reduction
  - ✓The gaining of electrons
  - ✓Decrease in the overall oxidation state

O is reduced (gains electrons).

$$\overset{-4}{C}\overset{+1}{H_4} + 2\,\overset{0}{O_2} \longrightarrow \overset{+4}{C}\overset{-2}{O_2} + 2\,\overset{+1}{H_2}\overset{-2}{O}$$

- Oxidation
  - ✓The loss of electrons
  - ✓Increase in the overall oxidation state



$$\overset{-4}{C}\overset{+1}{H_4} + 2\,\overset{0}{O_2} \longrightarrow \overset{+4}{C}\overset{-2}{O_2} + 2\,\overset{+1}{H_2}\overset{-2}{O}$$

C is oxidized (loses electrons).

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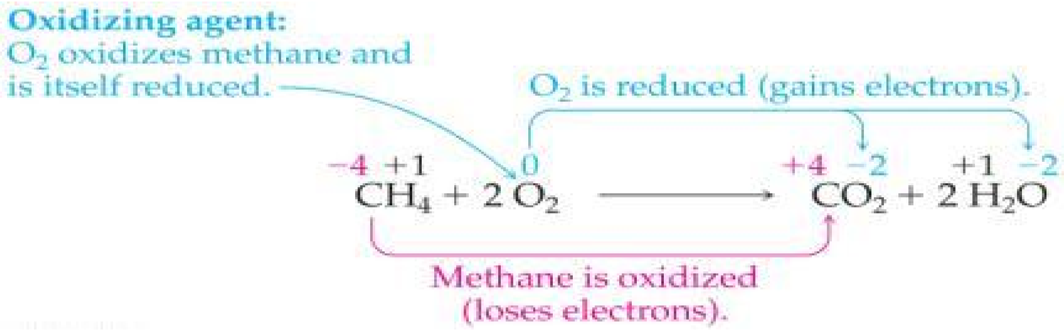
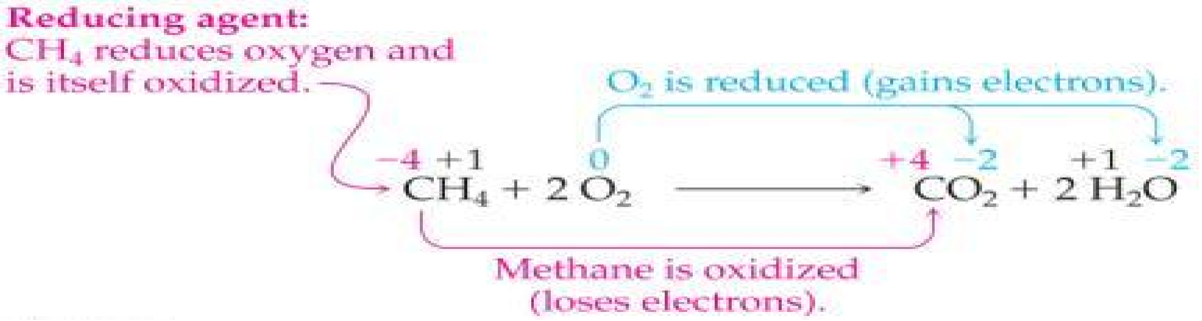
- **Oxidizing agent**
  - ✓Substance that oxidizes something else

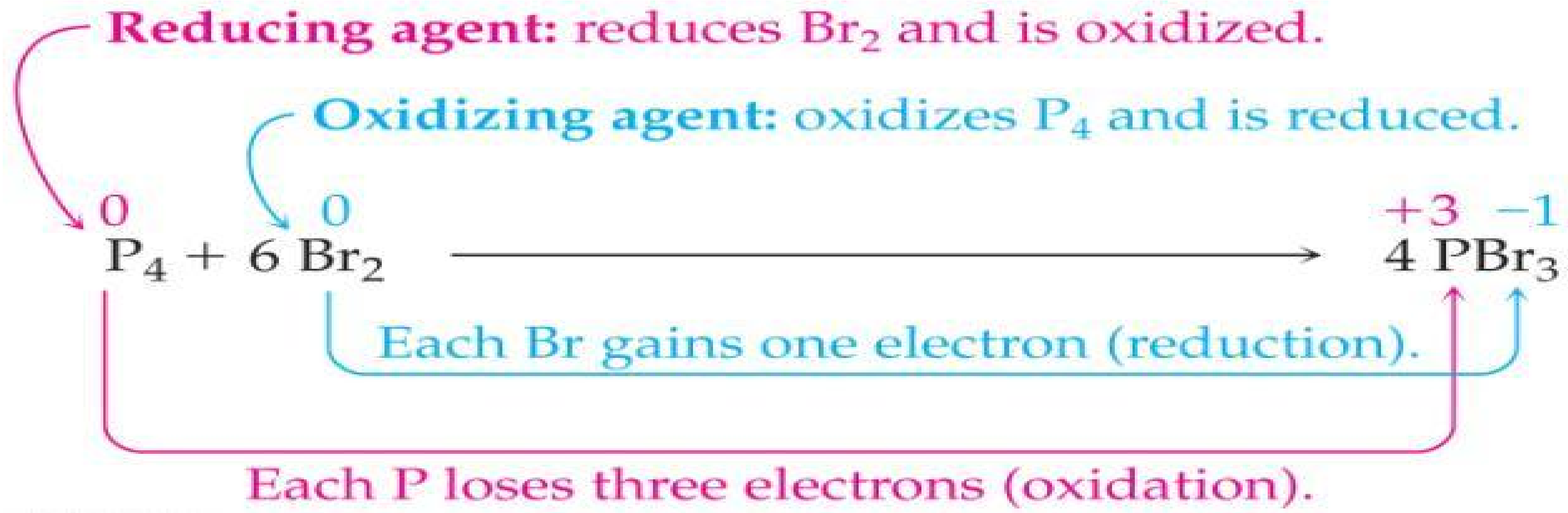


- **Reducing agent**
  - ✓Substance that reduces something else



Which substance is the oxidizing agent and which is the reducing agent in the following reaction?

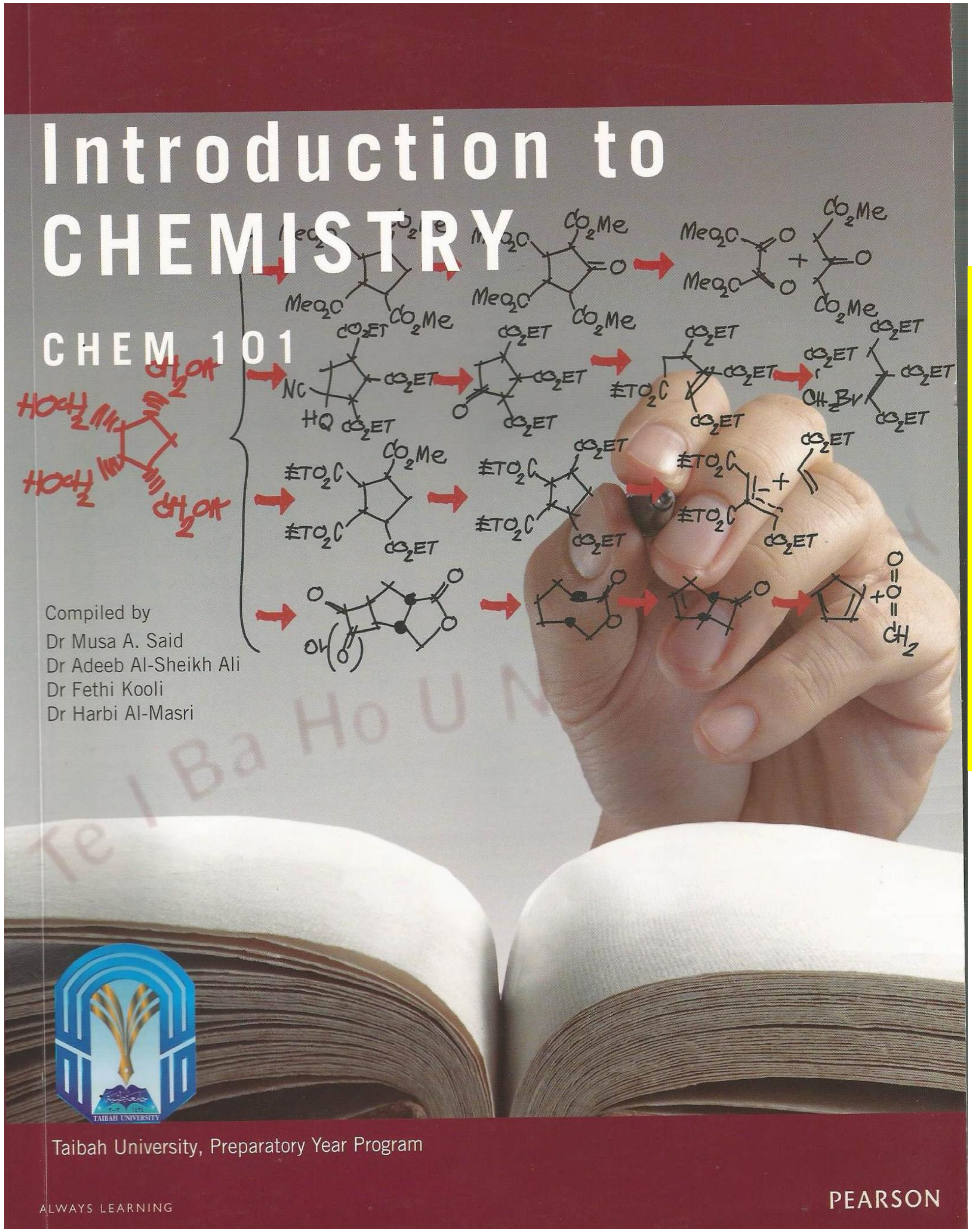
$$P_4 + 6\ Br_2 \rightarrow 4\ PBr_3$$



Review

In this les-son, we learned how to:

- Write equa-tions for acid-base re-ac-tions.
- Cal-cu-late ox-i-da-tion state.
- Iden-tify ox-i-dant and re-duc-tant using ox-i-da-tion states.



Introduction to
CHEMISTRY

CHEM 101

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Taibah University, Preparatory Year Program

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Chapter 4

CHEMICAL BONDING
CHEMICAL REACTION

TOPIC 15

Types of Bonding
Representing Valence Electrons with Dots

Introduction

OBJECTIVES

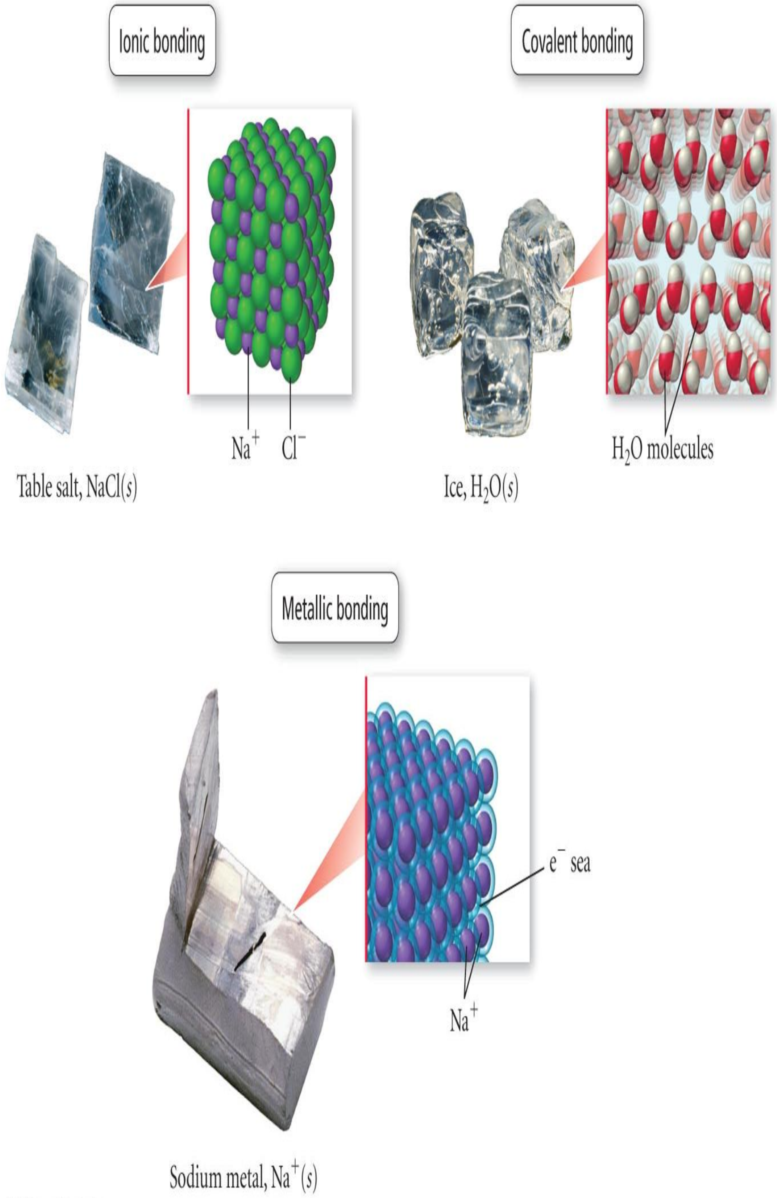
4.2.1 Identify chemical bonds and valence electrons

4.2.2 Identify and write Lewis dot structures

Chem-i-cal bonds may be ionic, co-va-lent, or metal-lic, with each type hav-ing char-ac-ter-is-tic prop-er-ties that re-sult from the trans-fer, shar-ing, or pool-ing of va-lence elec-trons in the outer en-ergy lev-els of the atoms.

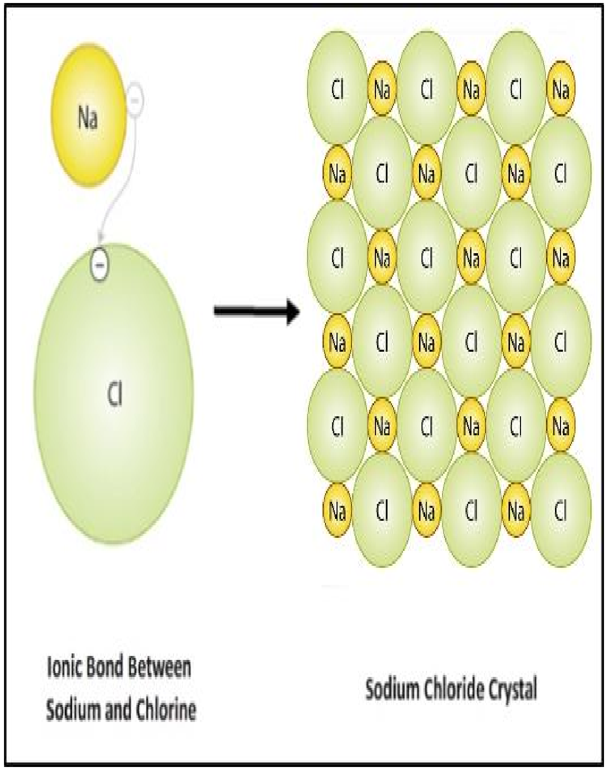
4.9 Types of Bonding

- **Three basic types of bonds**
  - **Ionic**
    - ➢Electrostatic attraction between ions.
  - **Covalent**
    - ➢Sharing of electrons.
  - **Metallic**
    - ➢Metal atoms bonded to several other atoms.



Ionic Bonds

- When a metal atom loses electrons it becomes a **cation**
  - ✓ metals have low ionization energy, making it *relatively* easy to remove electrons from them
- When a nonmetal atom gains electrons it becomes an **anion**
  - ✓ nonmetals have high electron affinities, making it advantageous to add electrons to these atoms
- The oppositely charged ions are then attracted to each other, resulting in an **ionic bond**



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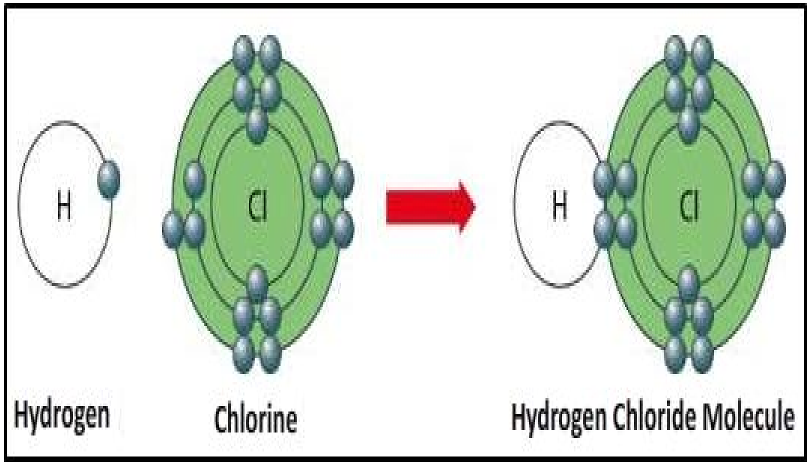
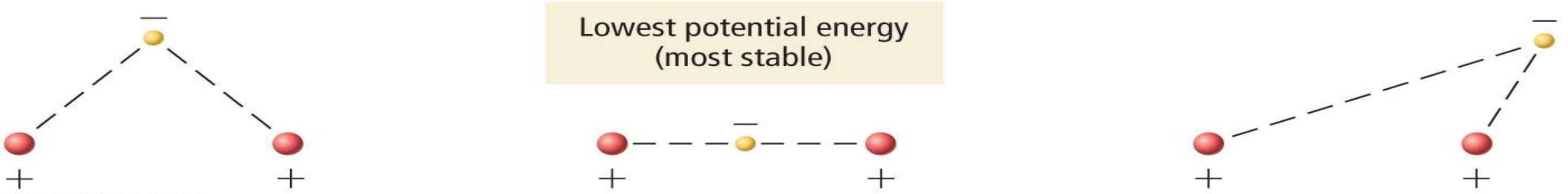
Covalent Bonds

- Nonmetal atoms have relatively high ionization energies, so it is difficult to remove electrons from them
- When nonmetals bond together, it is better in terms of potential energy for the atoms to share valence electrons
  - ✓ potential energy lowest when the electrons are between the nuclei
- Shared electrons hold the atoms together by attracting nuclei of both atoms





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Multiple Covalent Bonds

- **Single bond**—A covalent bond formed by sharing one electron pair.
  - ✓Represented by a single line: H-H
- **Double bond**—A covalent bond formed by sharing two electron pairs.
  - ✓Represented by a double line: O=O
- **Triple bond**—A covalent bond formed by sharing three electron pairs.
  - ✓Represented by a triple line: N≡N

Metallic Bonds

- The *relatively* low ionization energy of metals allows them to lose electrons easily
- The simplest theory of metallic bonding involves the metal atoms releasing their valence electrons to be shared as a pool by all the atoms/ions in the metal
  - ✓ an organization of metal cation islands in a sea of electrons
  - ✓ electrons delocalized throughout the metal structure
- Bonding results from attraction of cation for the delocalized electrons

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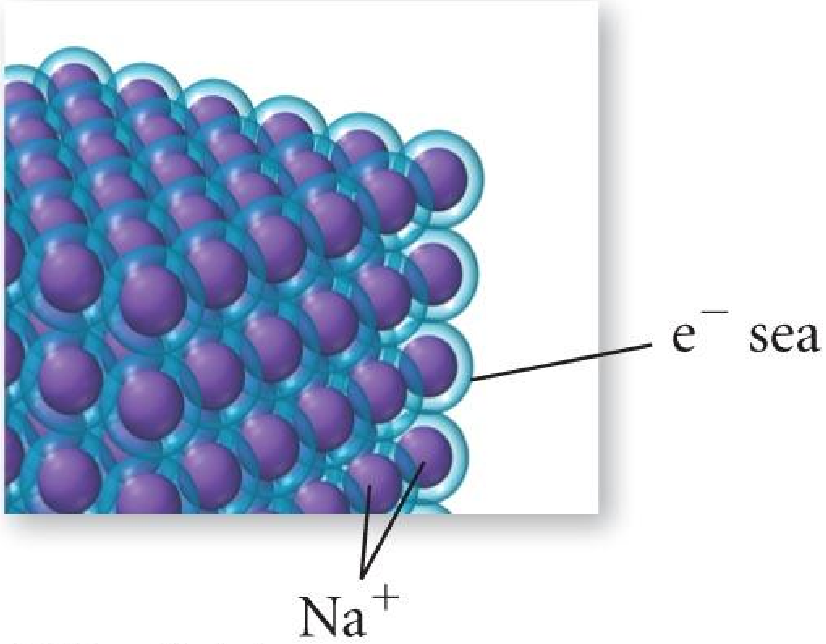
Metallic Bonding



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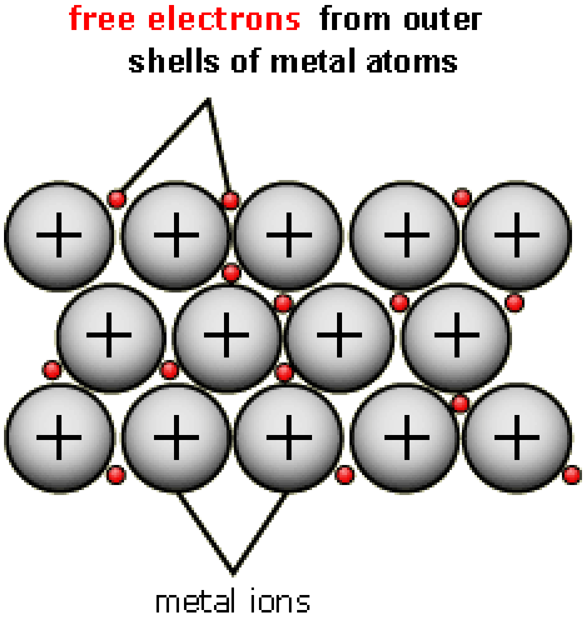
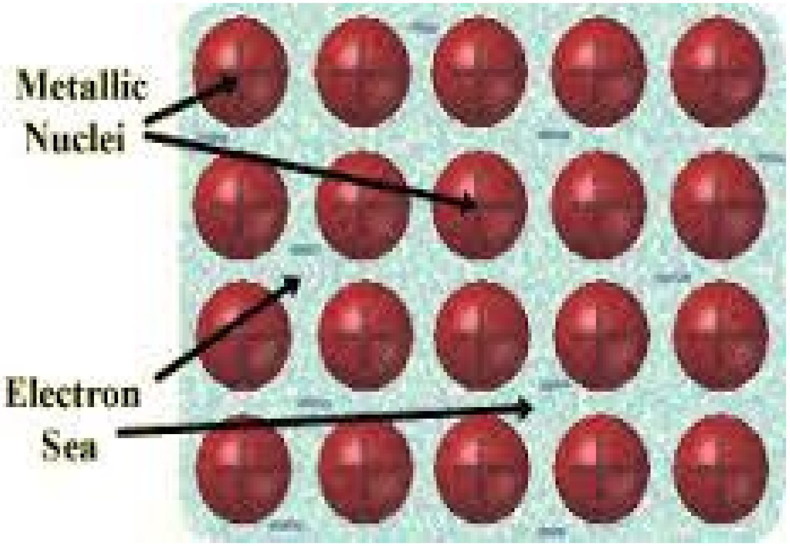
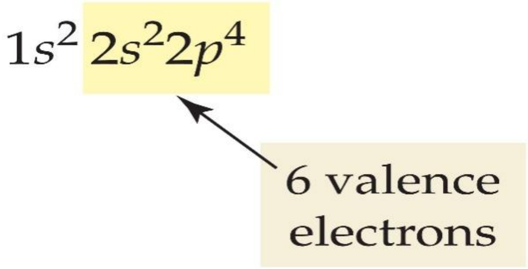




4.10 Representing Valence Electrons with Dots

- The electron configuration of O is as follows:



$1s^2\,2s^2 2p^4$

6 valence electrons

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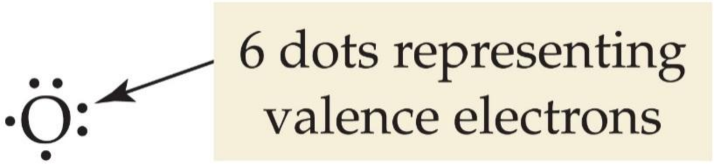
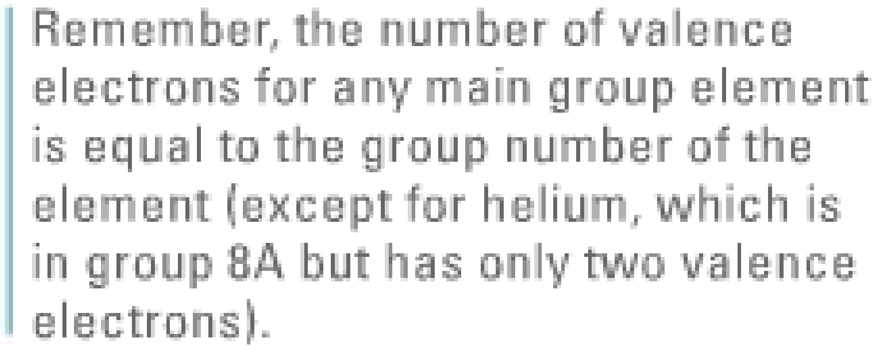
Its Lewis structure is as follows:



6 dots representing valence electrons



Remember, the number of valence electrons for any main group element is equal to the group number of the element (except for helium, which is in group 8A but has only two valence electrons).

Representing Valence Electrons with Dots

im_Representing Valence Electrons with Dots_23_03.mp

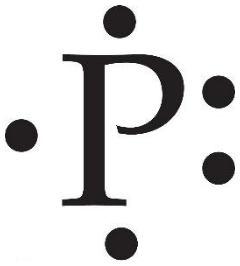
- Each dot represents a valence electron. The dots are placed around the element's symbol with a maximum of two dots per side.
- Although the exact location of dots is not critical, here we ***fill in the dots singly first and then pair them*** (with the exception of helium, described shortly).
- It is important to follow this convention when doing the homework.

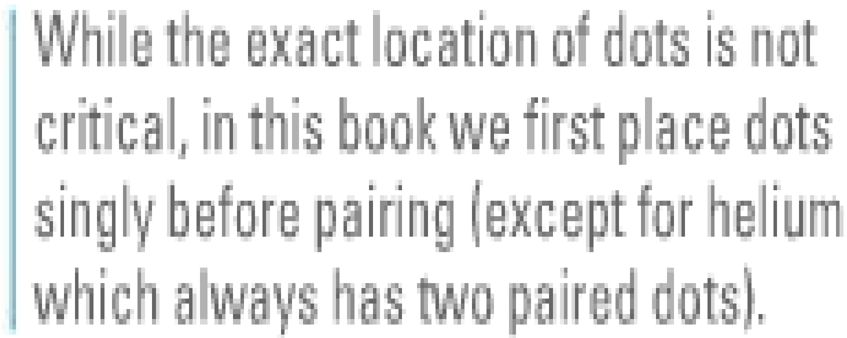
The Lewis Structures for All of the Period 2 Elements Are:

- Lewis structures allow us to easily see the number of valence electrons in an atom.
- Atoms with 8 valence electrons—which are particularly stable—are easily identified because they have eight dots, an **octet**.

Writing Lewis Structures for Elements

- Write the Lewis structure of phosphorus.
- Since phosphorus is in Group 5A in the periodic table, it has 5 valence electrons. Represent these as five dots surrounding the symbol for phosphorus.
- **Solution:**




While the exact location of dots is not critical, in this book we first place dots singly before pairing (except for helium which always has two paired dots).

Helium Is an exception. Its electron Configuration and Lewis Structure Are:

$1s^2$ He:

- The Lewis structure of helium contains two paired dots (a **duet**).
- For helium, a duet represents a stable electron configuration.

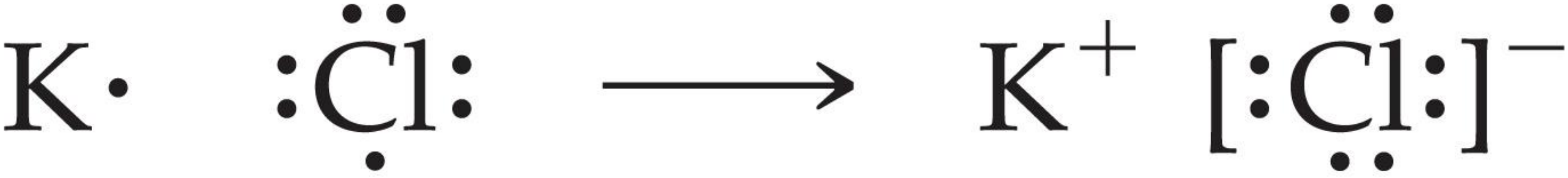
4.11 Ionic Bonding: Lewis Structures and Lattice Energies

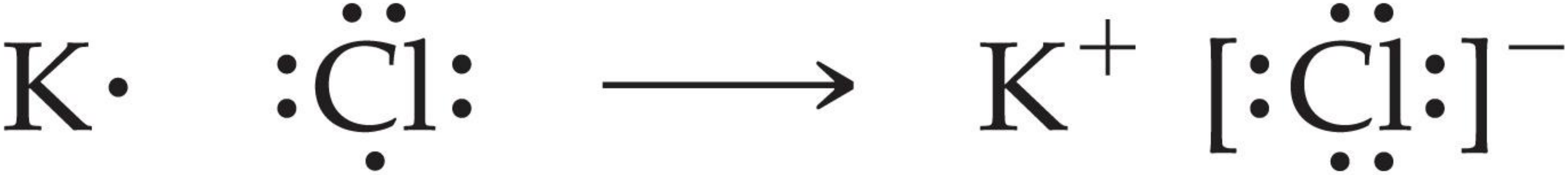
- When metals bond with nonmetals, electrons are transferred from the metal to the nonmetal.
- The metal becomes a cation and the nonmetal becomes an anion.
- The attraction between the cation and the anion results in an ionic compound.
- In Lewis theory, we represent this by moving electron dots from the metal to the nonmetal.

Lewis Structures of Ionic Compounds: Electrons Transferred

- For example, potassium and chlorine have the Lewis structures:

- When potassium and chlorine bond, potassium transfers its valence electron to chlorine.





$$K\cdot \quad :\underset{\cdot}{\overset{\cdot\cdot}{Cl}}: \longrightarrow K^{+}\ [:\underset{\cdot\cdot}{\overset{\cdot\cdot}{Cl}}:]^{-}$$

- The transfer of the electron gives chlorine an octet and leaves potassium with an octet in the previous principal shell, which is now the valence shell.
- The potassium, because it lost an electron, becomes positively charged, while the chlorine, which gained an electron, becomes negatively charged.
- The positive and negative charges attract one another, forming the compound KCl.
- The Lewis structure of an anion is usually written within brackets with the charge in the upper right corner (outside the brackets).
- The charges are shown, but the brackets are not shown in the homework drawing program online.

Lewis Theory Predicts the Correct Chemical Formulas for Ionic Compounds

- Consider the ionic compound formed between sodium and sulfur. The Lewis structures for sodium and sulfur are as follows:

- Sodium must lose one valence electron to obtain an octet (in the previous principal shell), while sulfur must gain two electrons to obtain an octet.
- The compound that forms between sodium and sulfur requires two sodium atoms to every one sulfur atom. The Lewis structure is as follows:

- The correct chemical formula is Na_2S.

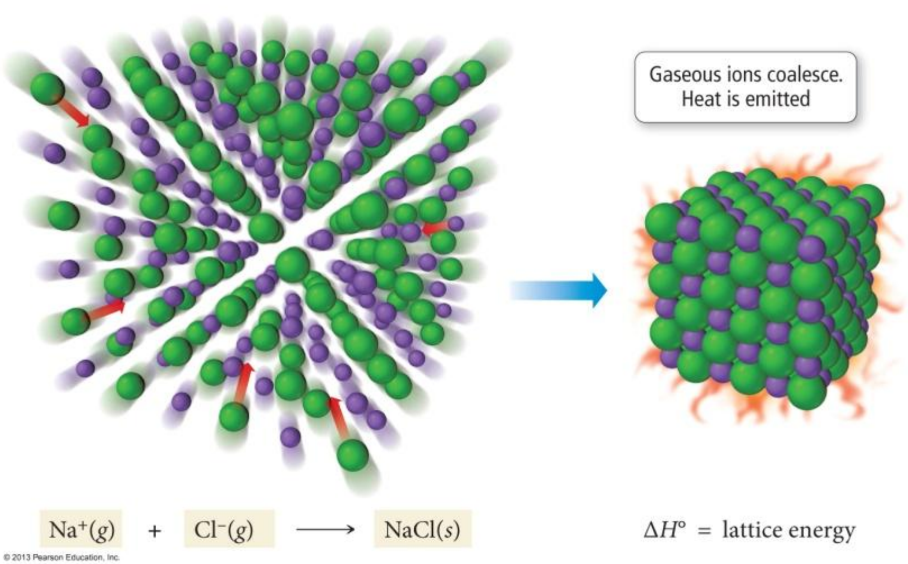
Lattice (Crystal) Energy

– The energy required to completely separate a mole of a solid ionic compound into its gaseous ions.

- The ions are arranged in a pattern called a crystal lattice.
  - Crystal lattice maximizes the attractions between cations and anions, leading to the most stable arrangement.

- The extra stability that accompanies the formation of the crystal lattice is measured as the lattice energy.

Lattice Energy of an Ionic Compound



Gaseous ions coalesce. Heat is emitted

$Na^+(g) + Cl^-(g) \longrightarrow NaCl(s)$ ΔH° = lattice energy

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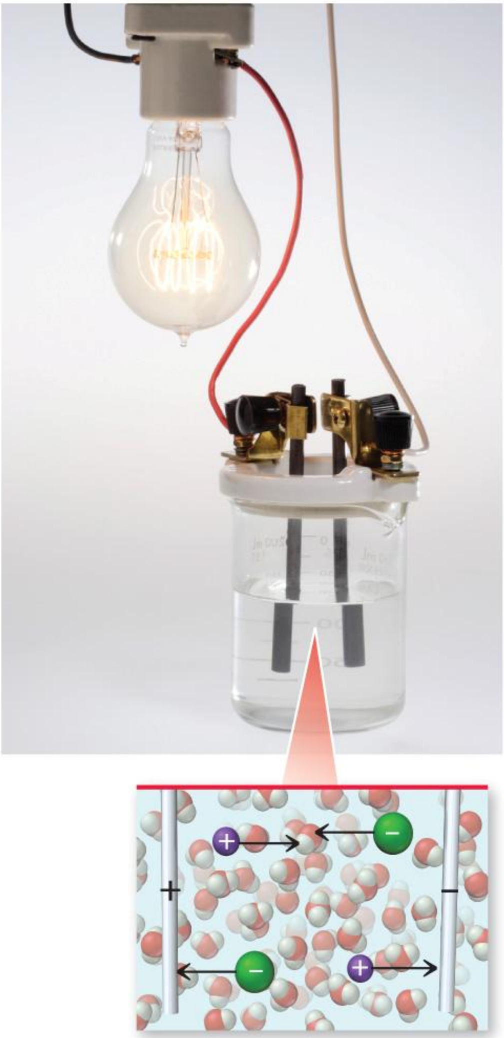
- Lattice energy increases with:
  - increasing charge on the ions
  - decreasing size of ions

Ionic Bonding: Models and Reality

- **Physical properties:**
  - ✓ **High melting points**
    - ➢ **Usually greater than 300 °C**
  - ✓ **High boiling points**
  - ✓ **Hard and brittle solids**

- **Solids:**
  - ✓ **All are crystalline solids at room temperature.**
  - ✓ **Do not conduct electricity, but when in their liquid state they act as strong electrolytes**
  - ✓ **Solid and liquid states are thermal insulators.**

- **Many solid ionic compounds are soluble in water.**
  - ✓ **Check solubility rules.**



NaCl(aq)

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Review

In this les-son, you learned how to:

- Iden-tify chem-i-cal bonds and va-lence elec-trons.
- Iden-tify and write Lewis dot struc-tures.

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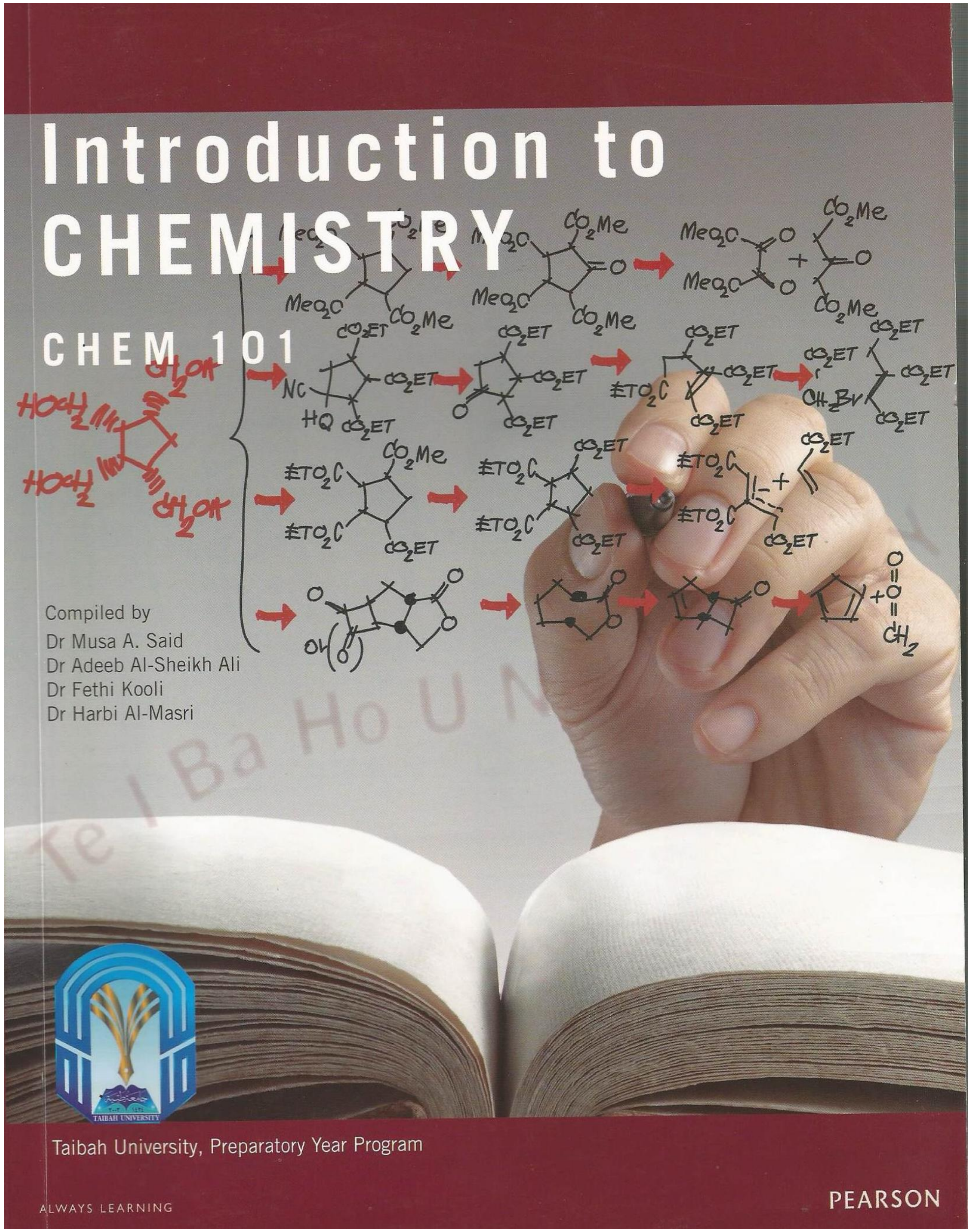

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Chapter 4

CHEMICAL BONDING
CHEMICAL REACTION

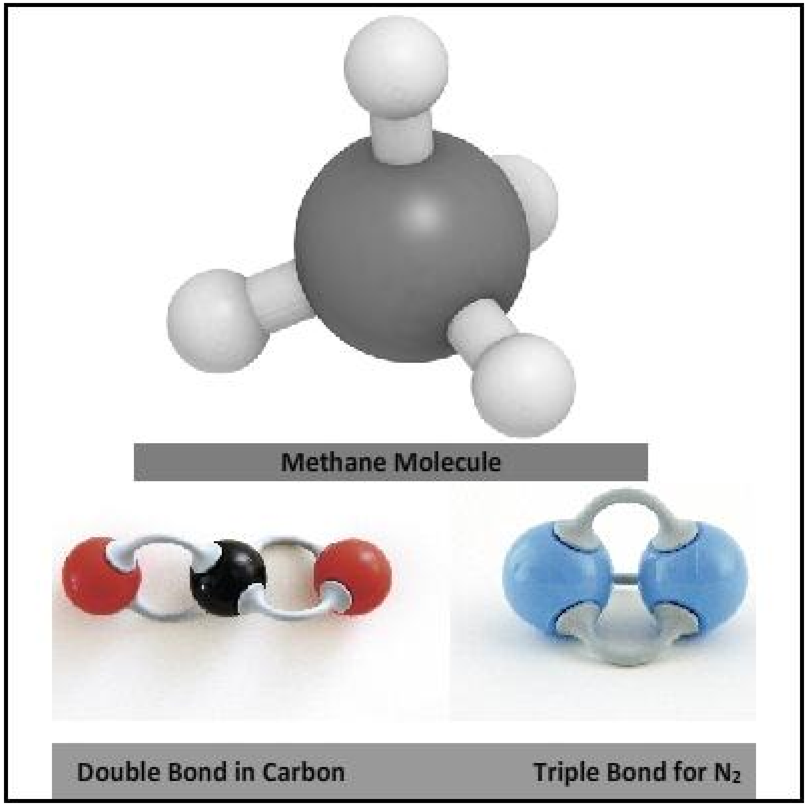
TOPIC 16

Lewis Structures

Electronegativity and Bond Polarity

Bond Energies and Bond Lengths

Introduction



OBJECTIVES

4.3.1 Classify bonds: pure covalent, polar covalent or ionic bond

4.3.2 Classify bonds according energy and length

The re-ac-tiv-ity of com-pounds not only de-pends upon the types of bonds, but also upon:

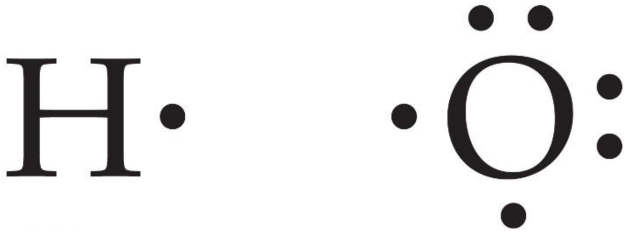
- The po-lar-ity of the mol-e-cules,
- The length be-tween the atoms, and
- The en-ergy re-quired to break the mol-e-cules apart.

4.12 Covalent Bonding: Lewis Structures

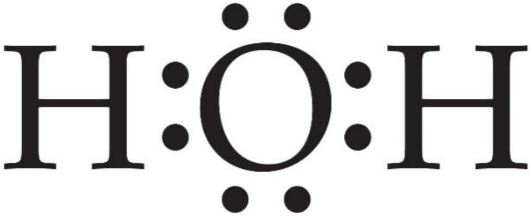
- When nonmetals bond with other nonmetals, a molecular compound results.
- Molecular compounds contain covalent bonds in which electrons are shared between atoms rather than transferred.
- In Lewis theory, we represent covalent bonding by allowing neighboring atoms to share some of their valence electrons in order to attain octets (or duets for hydrogen).

Covalent Lewis Structures: Electrons Shared

- Hydrogen and oxygen have the Lewis structures:

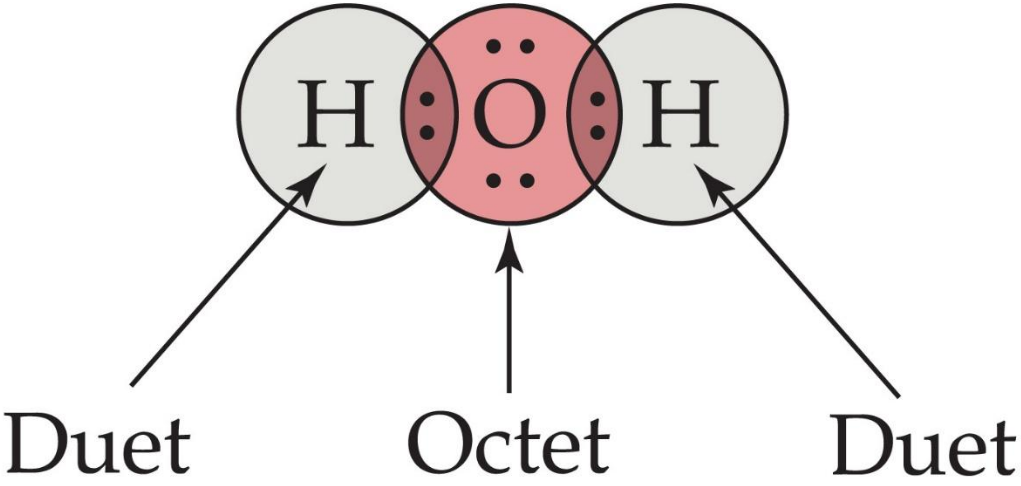


- In water, hydrogen and oxygen share their electrons so that each hydrogen atom gets a duet and the oxygen atom gets an octet.



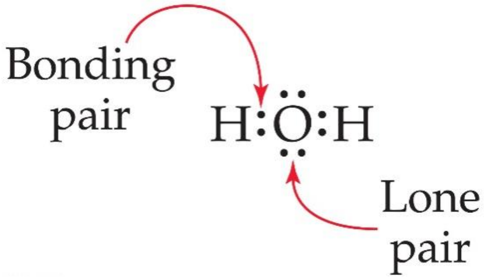
Covalent Lewis Structures: Electrons Shared

- The shared electrons—those that appear in the space between the two atoms—count toward the octets (or duets) of *both of the atoms*.



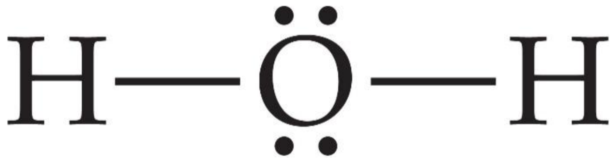
Electrons that are shared between two atoms are called ***bonding pair*** electrons.

Electrons that are only on one atom are called ***lone pair*** (or nonbonding) electrons.



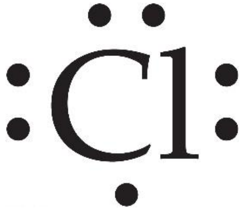
Bonding pair electrons are often represented by dashes to emphasize that they are a chemical bond.

Remember that each dash represents a *pair* of shared electrons.



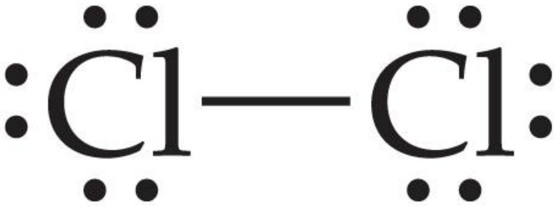
Lewis Theory Explains Why the Halogens Form Diatomic Molecules

- Consider the Lewis structure of chlorine.



- If two Cl atoms pair together, they can each attain an octet.

or



- When we examine elemental chlorine, we find that it exists as a diatomic molecule, just as Lewis theory predicts.
- The same is true for the other halogens.

Lewis Theory Predicts That Hydrogen Should Exist as H_2

- Hydrogen has the Lewis structure:

H·

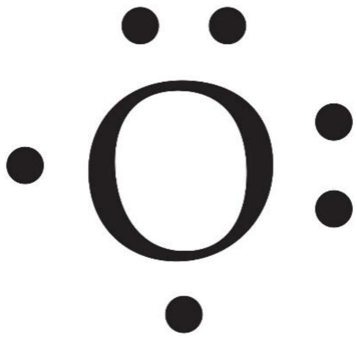
- When two hydrogen atoms share their valence electrons, they each get a duet, a stable configuration for hydrogen.

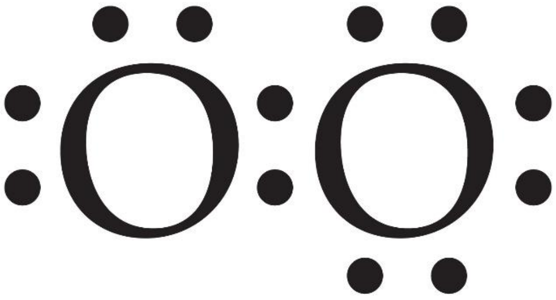
H:H *or* H—H

- Lewis theory is correct. In nature, elemental hydrogen exists as H_2 molecules.

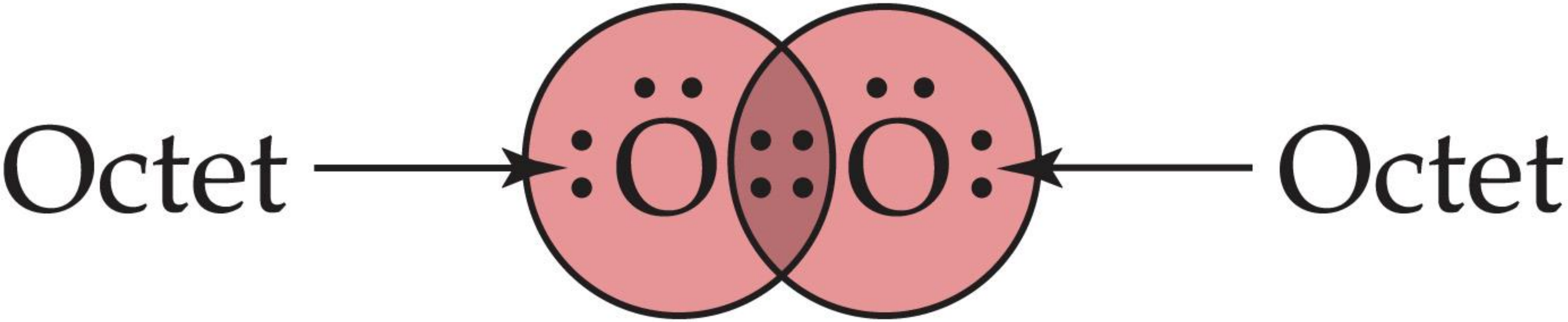
Double Bonds: In Lewis Theory, Atoms Can Share More Than One Electron Pair to Attain an Octet

- We know from Chapter 5 that oxygen exists as the diatomic molecule.
- If we pair two oxygen atoms together and then try to write a Lewis structure, we do not have enough electrons to give each O atom an octet.
- We can convert a lone pair into an additional bonding pair by moving it into the bonding region.



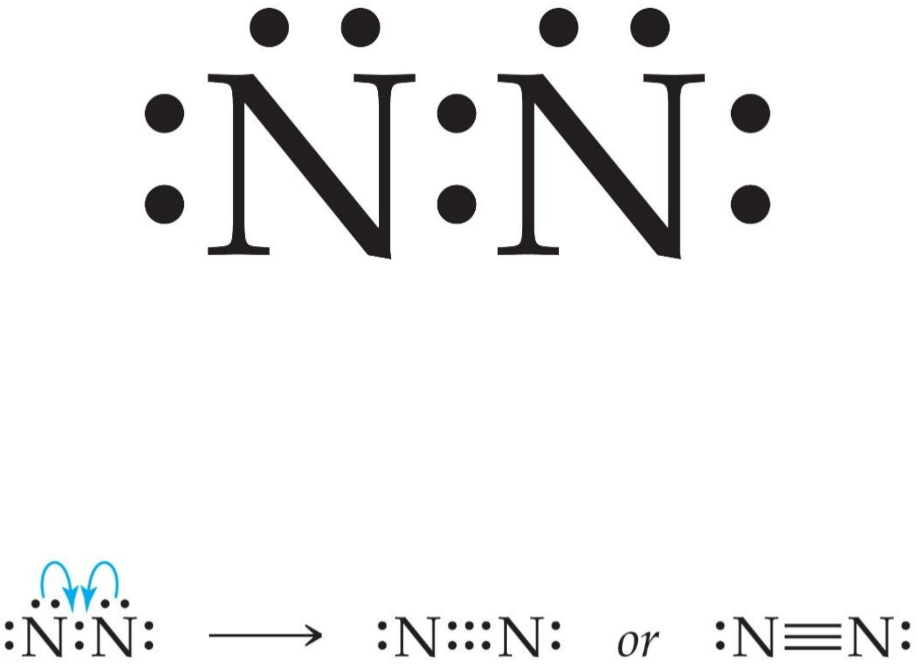


- Each oxygen atom now has an octet because the additional bonding pair counts toward the octet of both oxygen atoms.

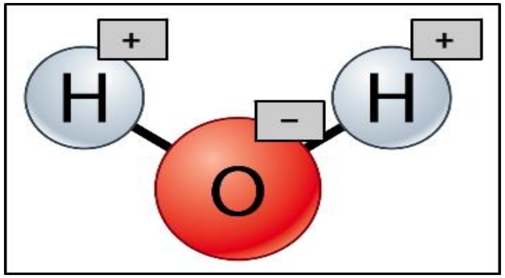


Triple Bonds: In Lewis Theory, Atoms Can Share More Than One Electron Pair to Attain an Octet

- Consider the Lewis structure of N_2.
- Since each N atom has 5 valence electrons, the Lewis structure for N_2 has 10 electrons.
- We do not have enough electrons to satisfy the octet rule for both N atoms.
- If we convert two lone pairs into bonding pairs, each nitrogen atom can get an octet.

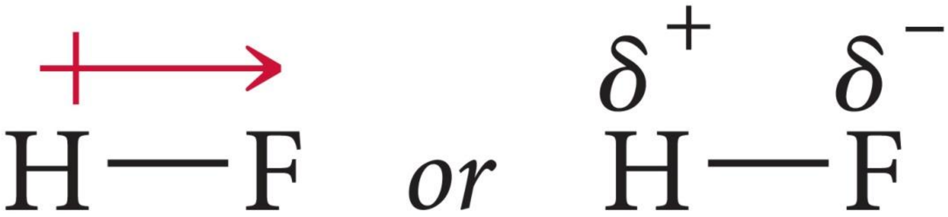


4.13 Electronegativity and Bond Polarity



Electronegativity

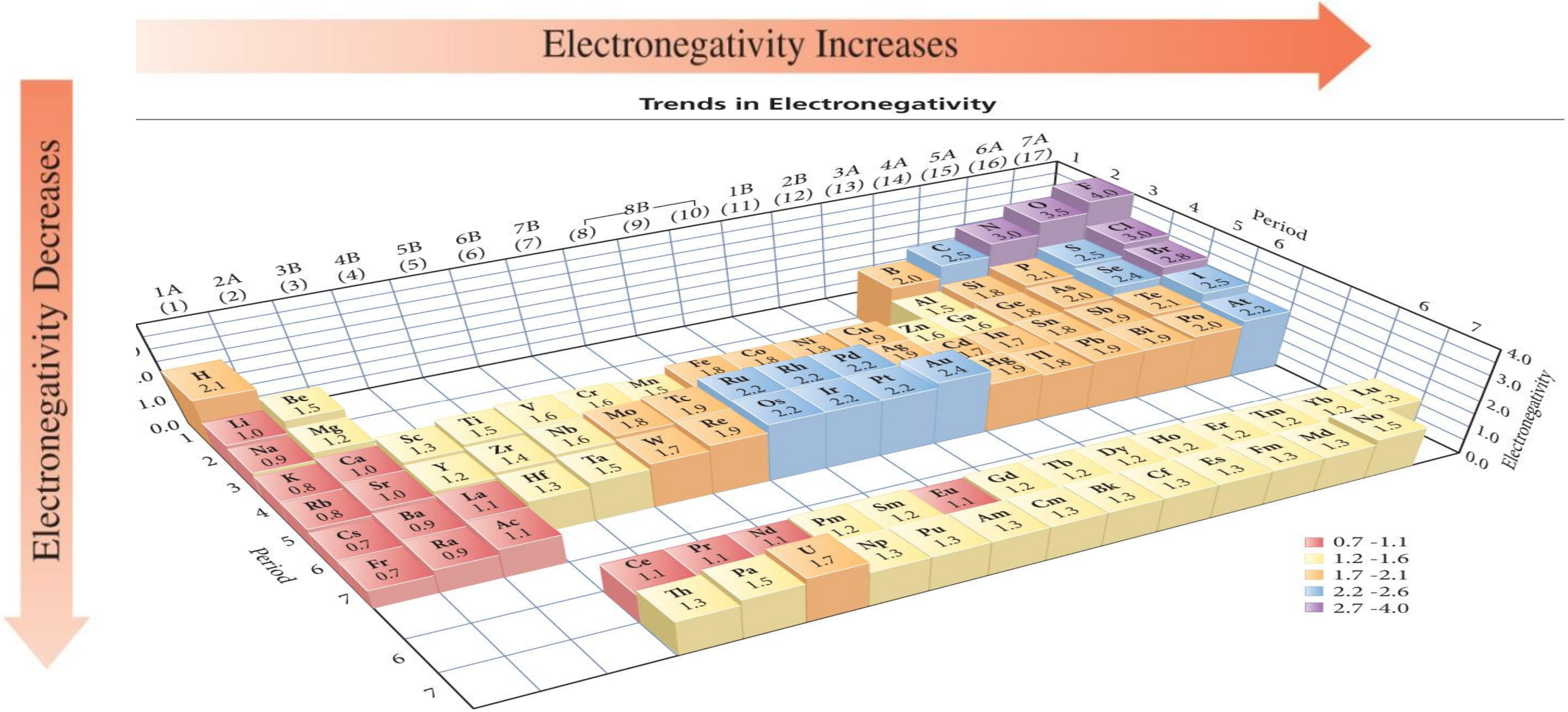
- is the relative ability of atoms to attract shared electrons
- is higher for nonmetals; fluorine has the highest with a value of 4.0
- is lower for metals; cesium and francium have the lowest value of 0.7



$\overset{\longmapsto}{\mathrm{H—F}}$ *or* $\overset{\delta^+}{\mathrm{H}}\text{—}\overset{\delta^-}{\mathrm{F}}$

▪**The greater the difference, the more POLAR the bond.**

Electronegativity Values and Trend



Trend:

- **Electronegativity generally increases across a period in the periodic table.**
- **Electronegativity generally decreases down a column in the periodic table.**
- **Fluorine is the most electronegative element.**
- **Francium is the least electronegative element.**

Electronegativity and Bond Polarity

If the difference in electronegativity between bonded atoms is:

- ✓ **Zero (0), then the bond is pure covalent.**
  - ➢ **Equal sharing of electrons between atoms**

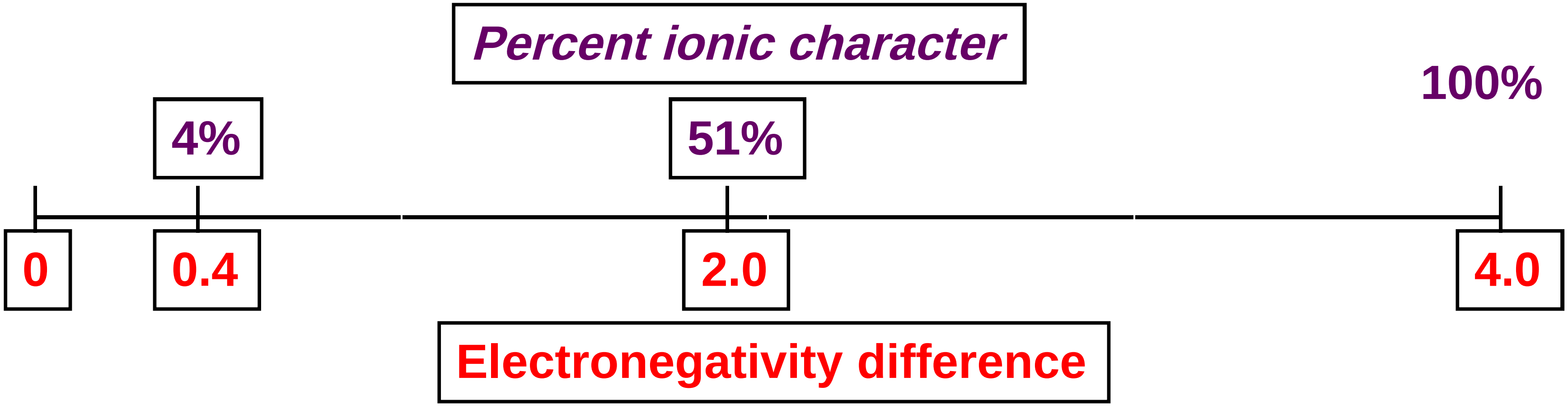
anim_polar covalent_24_02.mp4

- ✓ **0.1 to 0.4, then the bond is nonpolar covalent.**

- ✓ **0.5 to 1.9, then the bond is polar covalent.**

anim_Pure Covalent_24_01.mp4

- ✓ **≥2.0, then the bond is ionic.**
  - ➢ **Transfer of electrons between atoms**



Percent ionic character

4%

51%

100%

0

0.4

2.0

4.0

Electronegativity difference

Electronegativity and Bond Types

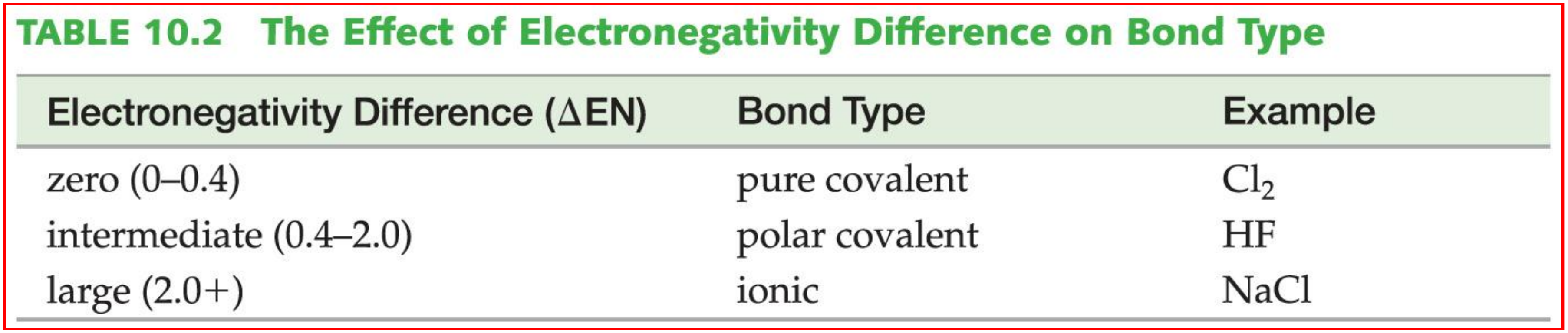
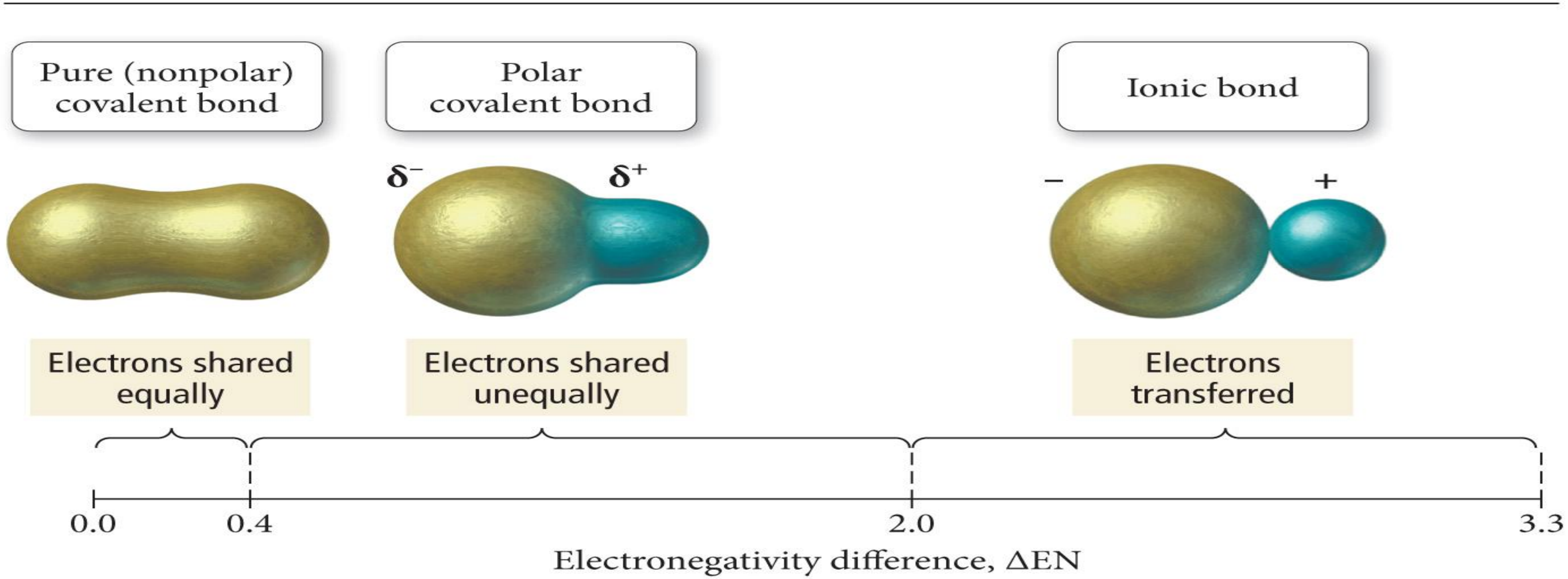


TABLE 10.2 The Effect of Electronegativity Difference on Bond Type

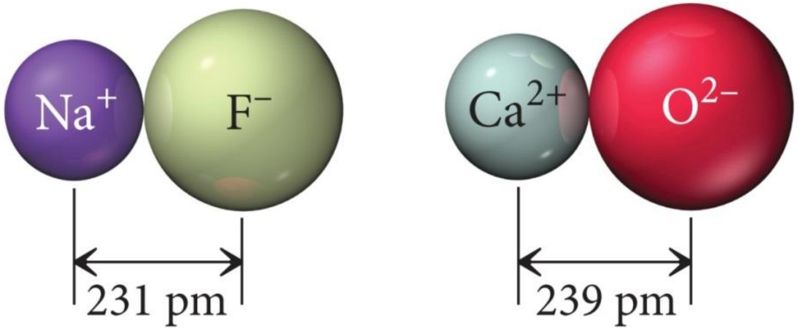
Electronegativity Difference (ΔEN)	Bond Type	Example
zero (0–0.4)	pure covalent	Cl_2
intermediate (0.4–2.0)	polar covalent	HF
large (2.0+)	ionic	NaCl

The Continuum of Bond Types



4.14 Bond Energies and Bond Lengths

- Chemical reactions involve breaking bonds in reactant molecules and making new bonds to create the products.
- The amount of energy, in the gaseous state, that it takes to break one mole of a bond in a compound is called the **bond energy**.



anim_Bond Length_24_03.mp4

Trends in Bond Energies

- In general, the more electrons two atoms share, the stronger the covalent bond.
  - Must be comparing bonds between like atoms
  - **C≡C (837 kJ) > C═C (611 kJ) > C—C (347 kJ)**
  - **C≡N (891 kJ) > C ═ N (615 kJ) > C—N (305 kJ)**

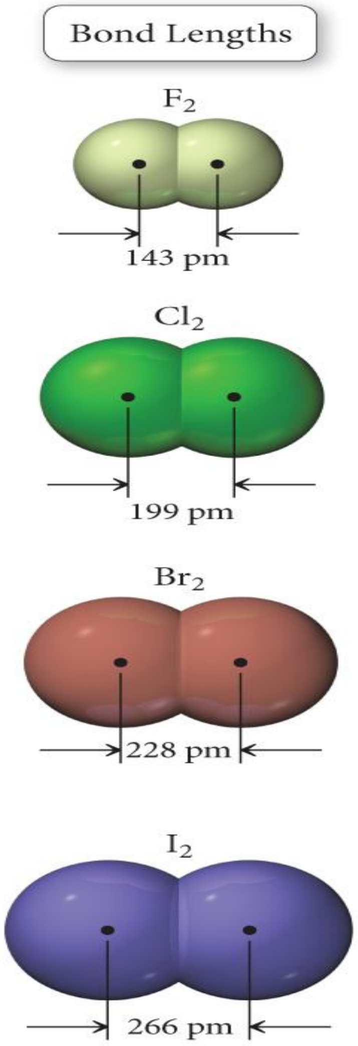
- In general, the shorter the covalent bond, the stronger the bond.
  - Must be comparing similar types of bonds
  - Br—F (237 kJ) > Br—Cl (218 kJ) > Br—Br (193 kJ)
  - Bonds get weaker down the column.
  - Bonds get stronger across the period.

Covalent Bonding: Model versus Reality for Bond Length

- Lewis theory predicts that the more electrons two atoms share, the shorter the bond should be.
  - ✓ When comparing bonds to like atoms

- **Bond length** is determined by measuring the distance between the nuclei of bonded atoms.

- In general, triple bonds are shorter than double bonds, and double bonds are shorter than single bonds.



Bond Lengths

- The distance between the nuclei of bonded atoms is called the **bond length**.

- Because the actual bond length depends on the other atoms around the bond, we often use the **average bond length**.
  - ✓ Averaged for similar bonds from many compounds

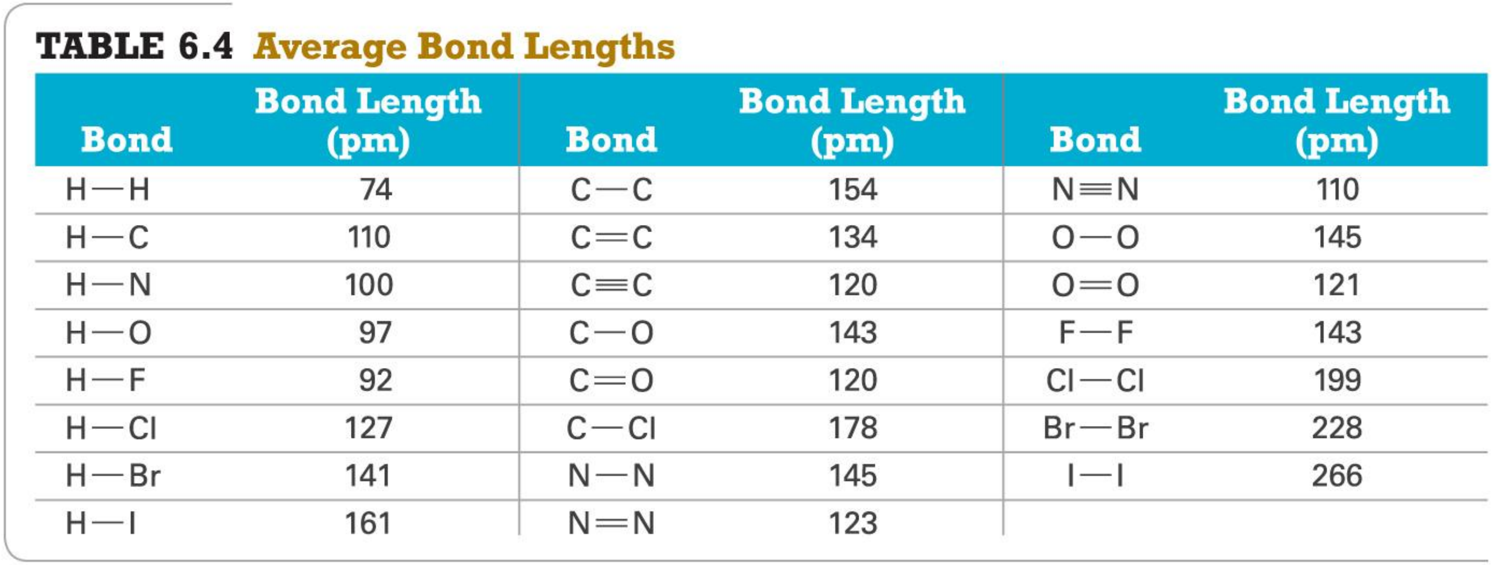


TABLE 6.4 **Average Bond Lengths**

Bond	Bond Length (pm)	Bond	Bond Length (pm)	Bond	Bond Length (pm)
H—H	74	C—C	154	N≡N	110
H—C	110	C═C	134	O—O	145
H—N	100	C≡C	120	O═O	121
H—O	97	C—O	143	F—F	143
H—F	92	C═O	120	Cl—Cl	199
H—Cl	127	C—Cl	178	Br—Br	228
H—Br	141	N—N	145	I—I	266
H—I	161	N═N	123		

Trends in Bond Lengths

- In general, the more electrons two atoms share, the shorter the covalent bond.
  - ✓ Must be comparing bonds between like atoms
  - ✓ C≡C (120 pm) < C═C (134 pm) < C—C (154 pm)
  - ✓ C≡N (116 pm) < C═N (128 pm) < C—N (147 pm)

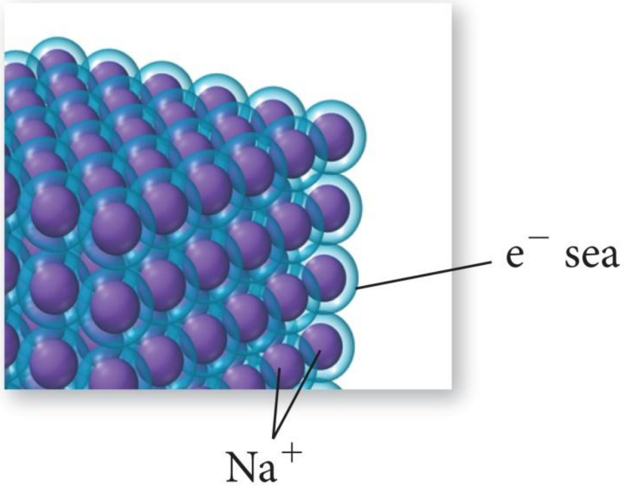
- **Generally, bond length decreases from left to right across a period.**
  - ✓ C—C (154 pm) > C—N (147 pm) > C—O (143 pm)

- **Generally, bond length increases down the column.**
  - ✓ F—F (144 pm) > Cl—Cl (198 pm) > Br—Br (228 pm)

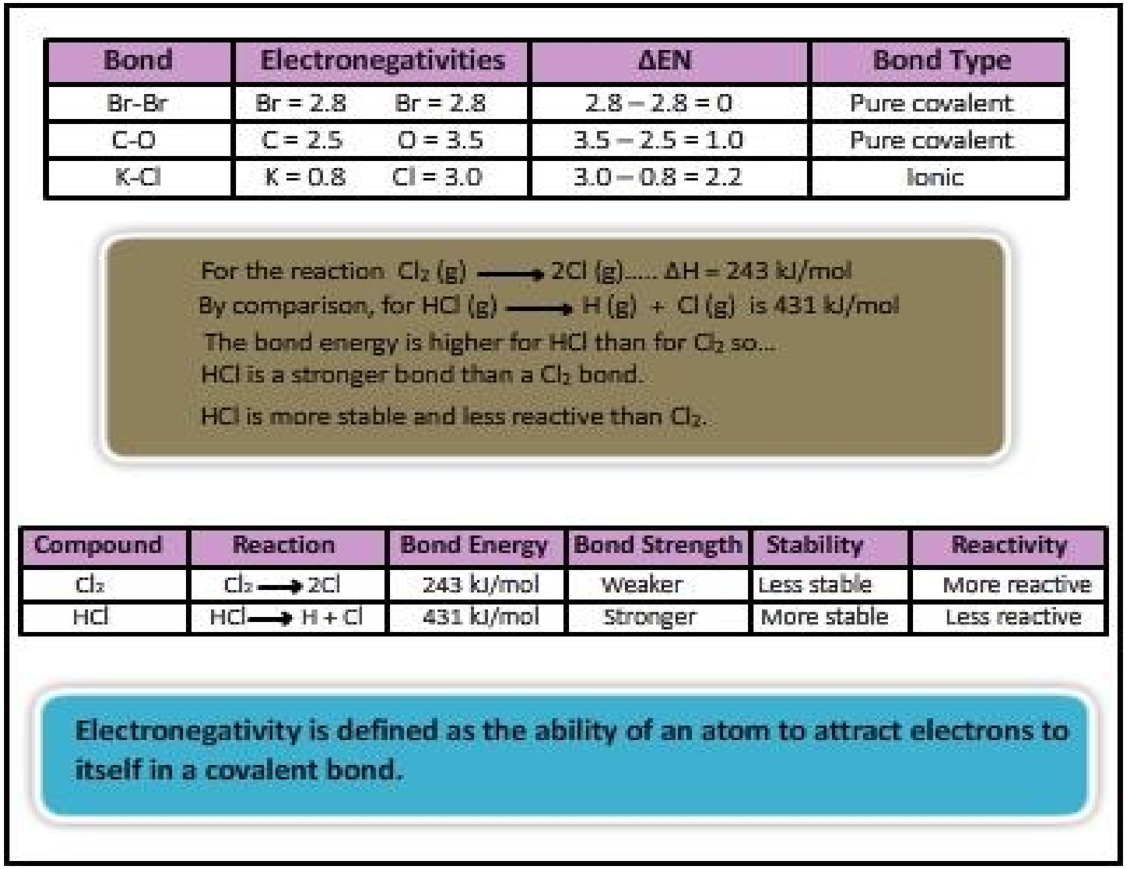
- In general, as bonds get longer, they also get weaker.

4.15 Bonding in Metals: The Electron Sea Model

- The simplest theory of metallic bonding involves the metal atoms releasing their valence electrons to be shared as a pool by all the atoms/ions in the metal.
  - ✓ An organization of metal cation islands in a sea of electrons
  - ✓ Electrons delocalized throughout the metal structure
- Bonding results from attraction of cation for the delocalized electrons.



Review



Bond	Electronegativities		ΔEN	Bond Type
Br-Br	Br = 2.8	Br = 2.8	2.8 – 2.8 = 0	Pure covalent
C-O	C = 2.5	O = 3.5	3.5 – 2.5 = 1.0	Pure covalent
K-Cl	K = 0.8	Cl = 3.0	3.0 – 0.8 = 2.2	Ionic

For the reaction $Cl_2(g) \longrightarrow 2Cl(g)$..... $\Delta H = 243$ kJ/mol
By comparison, for $HCl(g) \longrightarrow H(g) + Cl(g)$ is 431 kJ/mol
The bond energy is higher for HCl than for Cl_2 so...
HCl is a stronger bond than a Cl_2 bond.
HCl is more stable and less reactive than Cl_2.

Compound	Reaction	Bond Energy	Bond Strength	Stability	Reactivity
Cl_2	$Cl_2 \longrightarrow 2Cl$	243 kJ/mol	Weaker	Less stable	More reactive
HCl	$HCl \longrightarrow H + Cl$	431 kJ/mol	Stronger	More stable	Less reactive

Electronegativity is defined as the ability of an atom to attract electrons to itself in a covalent bond.

In this lesson, we have learned how to:

- Classify bonds: pure covalent, polar covalent, or ionic bond.
- Classify bonds according to energy and length.