

Chapter 4

Chemical Quantities and Aqueous Reactions

Read/Study: Chapter 4

Memorization: Work on 3 x 5 Cards

MGC Assignment: See MGC Website!

“I feel sorry for people who don’t understand anything about chemistry. They are missing an important source of happiness. – Linus Pauling

Chapter 4

Chemical Quantities and Aqueous Reactions

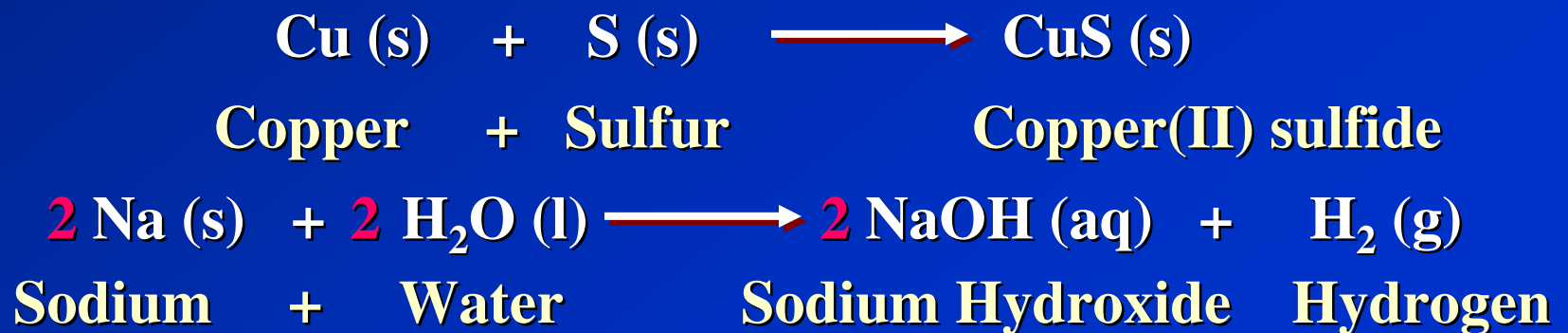
1. Writing and Balancing Equations
2. Reaction Stoichiometry
 - A. The Process
 - B. Limiting Reagent Problems
 - C. Calculating Yields
3. Reactions in Water

“I feel sorry for people who don’t understand anything about chemistry. They are missing an important source of happiness. – Linus Pauling

Chemical Equations - The sentences of the chemical language.

Reactants: The original or starting substances in a chemical reaction; normally written on the *left* side of the chemical equation.

Products: The new or formed substances in a chemical reaction; normally written on the *right* side of the chemical equation.



The “Meaning” of a Chemical Equation:

1 Molecule

2 molecules

1 molecule

2 molecules

1 Mole

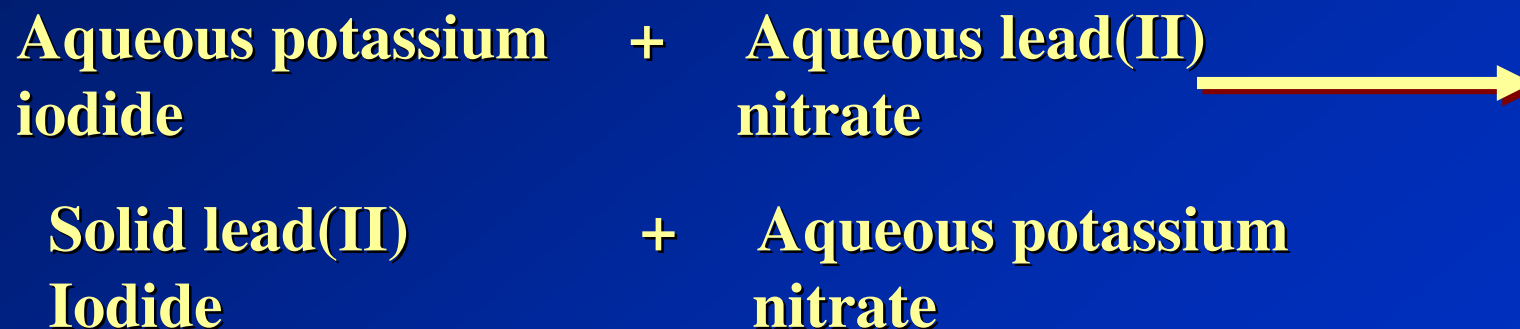
2 Moles

1 Mole

2 Moles

Rules for Writing Chemical Equations -

1. Identify reactants and products and write a “word equation.



2. Write symbols and formulas for the elements and compounds



3. Balance by changing the coefficients in front of the symbols and formulas. Do **NOT** change subscripts in the formulas or add or remove substances.



4. Check to see that the same number of each kind of atom is shown on both sides.



5. Add symbols showing the physical state of each substance; solid (s), liquid (l), gas (g) or aqueous solution (aq).

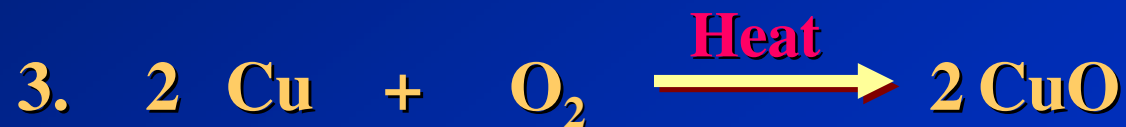


Practice Problem:

Write the balanced chemical equation for the reaction of hot copper metal with oxygen to form solid copper(II) oxide.



Solid copper(II) oxide



4. 2 copper atoms and 2 oxygen atoms on each side!!



Practice Problem:

Burn liquid heptane, C_7H_{16} , in oxygen gas to form carbon dioxide gas and liquid water.

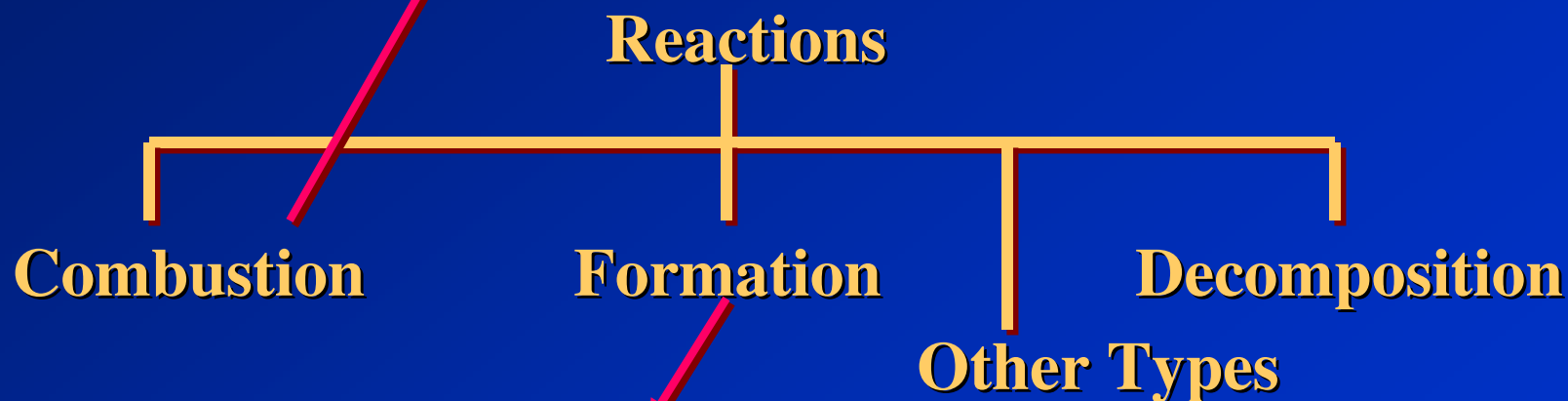


3. 7 Carbons, 16 hydrogens, and 22 oxygens on each side of the equation.



Introduction to the Classification of Reactions:

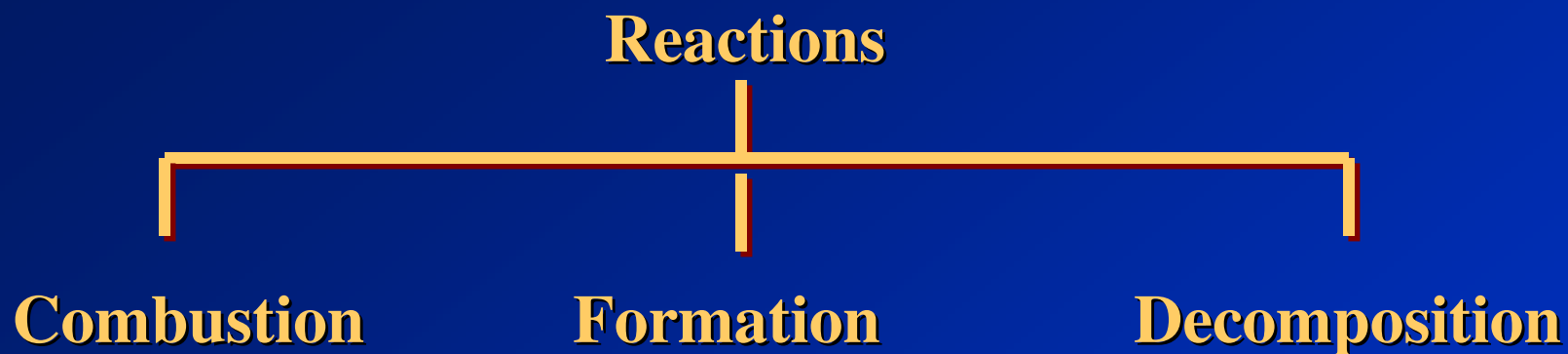
The last reaction was a “combustion reaction”. This type of reaction normally involves a rapid reaction of some substance (an element or compound) with oxygen.



Two or more elements chemically combine to form one compound.



Introduction to the Classification of Reactions:



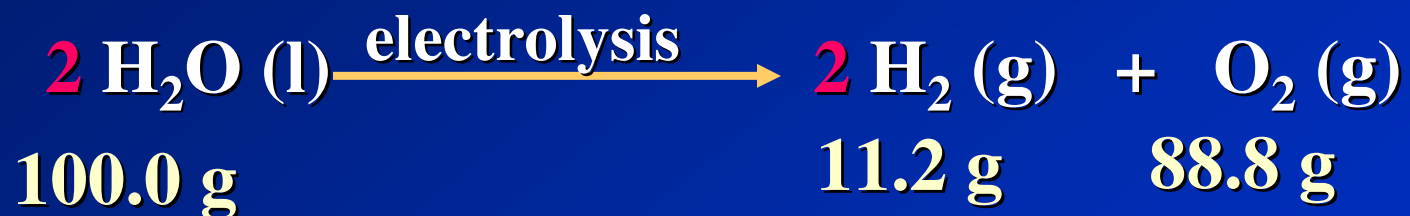
A single compound breaks apart to form two or more simpler substances; either simpler compounds or elements.



Important Law Governing Chemical Reactions

The Law of Conservation of Mass - There is no measurable change in total mass during a chemical reaction.

“Mass is neither created nor destroyed”



The Law of Conservation of Mass



Summary of Laws Governing Chemical Composition and Reactions

The Law of Conservation of Mass

The Law of Constant Composition

The Law of Definite Proportions

The Law of Multiple Proportions

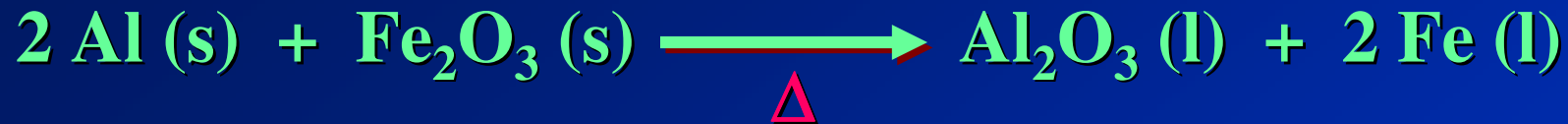
These laws that govern chemical reactions and compound composition allow us to do stoichiometric calculations. They were all well known by 1803 when John Dalton put forth his atomic theory of matter. This theory was able to “explain” why these laws that govern chemical reactions were true.

People Who Helped Us to Understand More About Chemistry

Marie & Antoine Lavoisier

John Dalton

Interpreting Chemical Equations:



2 atoms	1 formula unit	1 formula unit	2 atoms
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2 moles	1 mole	1 mole	2 moles
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53.964 g	159.69 g	101.96 g	111.69 g
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26.982 g	79.845 g	50.980 g	55.845 g
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Chemical equations express the quantitative relationships among the reactants and products in a chemical reaction.

The “Meaning” of a Chemical Equation:

1 Molecule

2 molecules

1 molecule

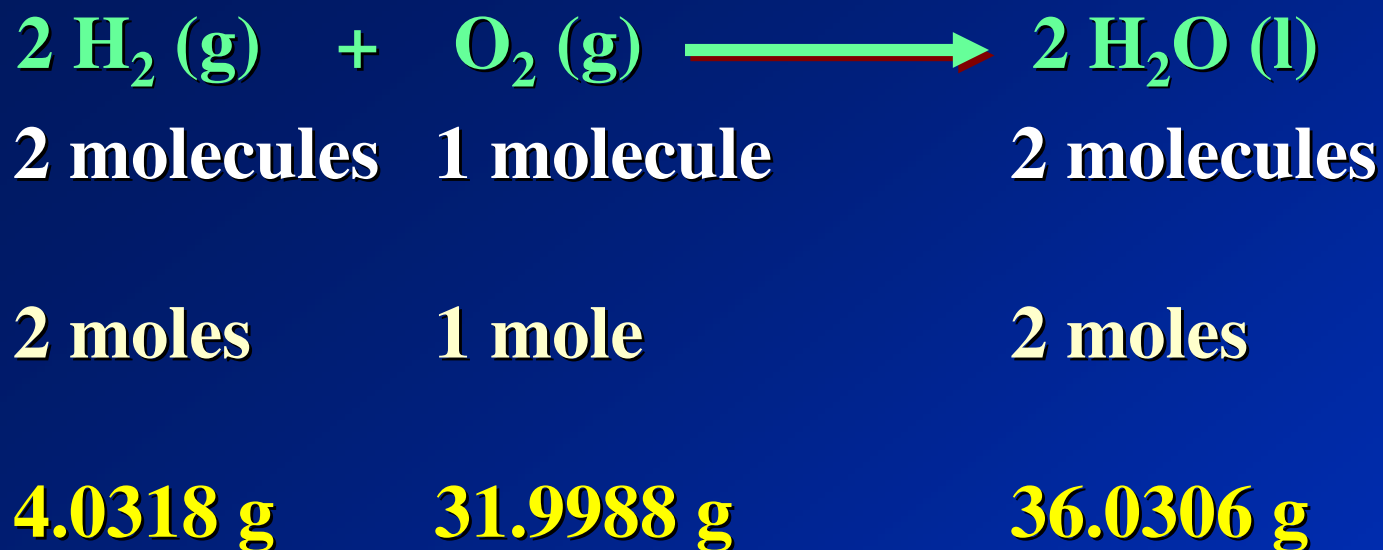
2 molecules

1 Mole

2 Moles

1 Mole

2 Moles



Mass Relationships in Chemical Reactions:

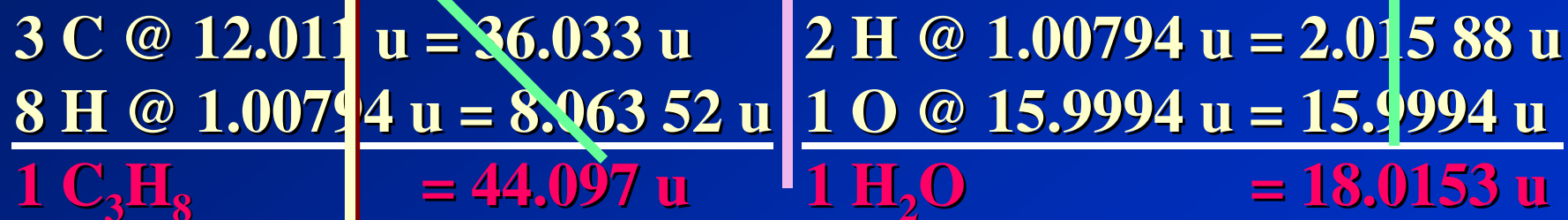
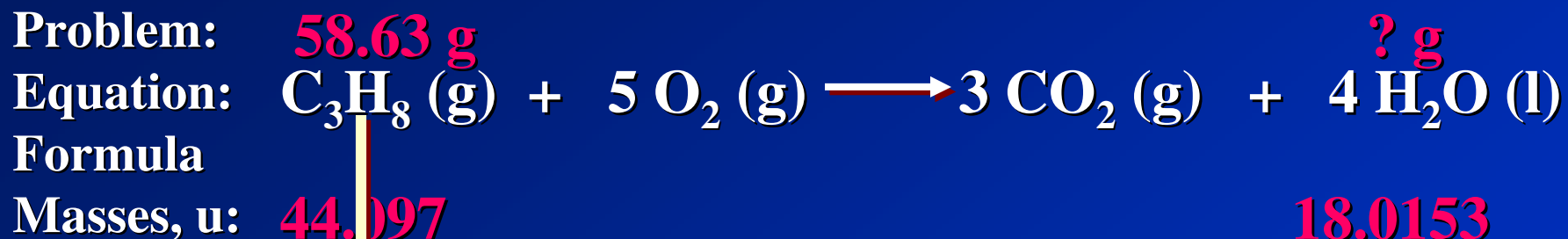
Calculate the number of grams of water produced when 58.63 grams of propane are burned.

- Write the balanced chemical equation and identify the known and unknown quantities:



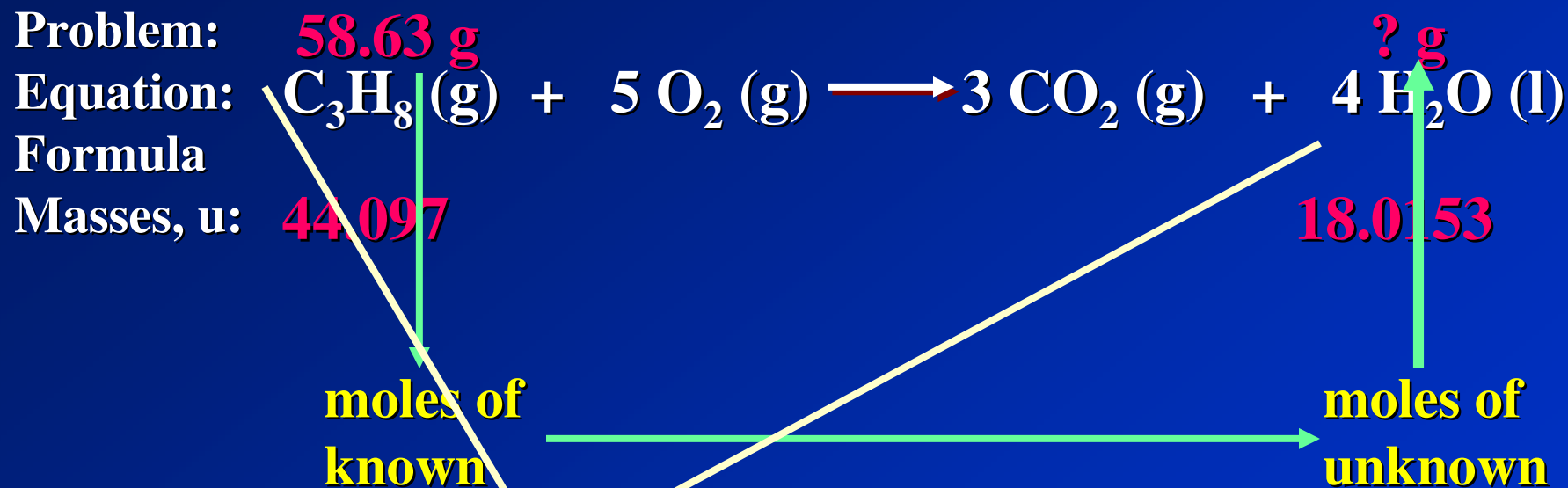
2. Convert the known quantities to moles:

(“When in doubt, calculate moles”)



$$\frac{(58.63 \text{ g } \text{C}_3\text{H}_8) (1 \text{ mol } \text{C}_3\text{H}_8)}{(44.097 \text{ g } \text{C}_3\text{H}_8)} = 1.3296 \text{ mol } \text{C}_3\text{H}_8$$

3. Use the mole ratio in the equation to find the moles of the unknown material:

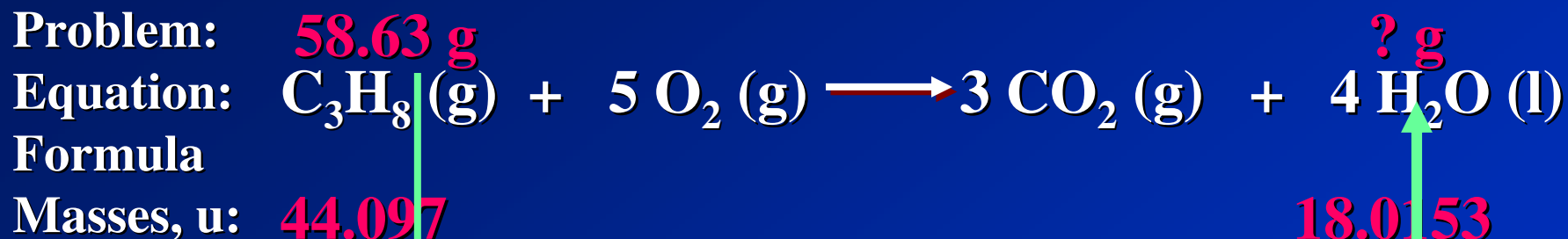


$$\cancel{(1.3296 \text{ mol C}_3\text{H}_8)} \times \frac{(4 \text{ mol H}_2\text{O})}{\cancel{(1 \text{ mol C}_3\text{H}_8)}} = 5.3184 \text{ mol H}_2\text{O}$$

4. Convert moles of the unknown material to the units called for in the problem:

$$\cancel{(5.3184 \text{ mol H}_2\text{O})} \times \frac{(18.0153 \text{ g H}_2\text{O})}{\cancel{(1 \text{ mol H}_2\text{O})}} = \boxed{95.81 \text{ g H}_2\text{O}}$$

Summary:



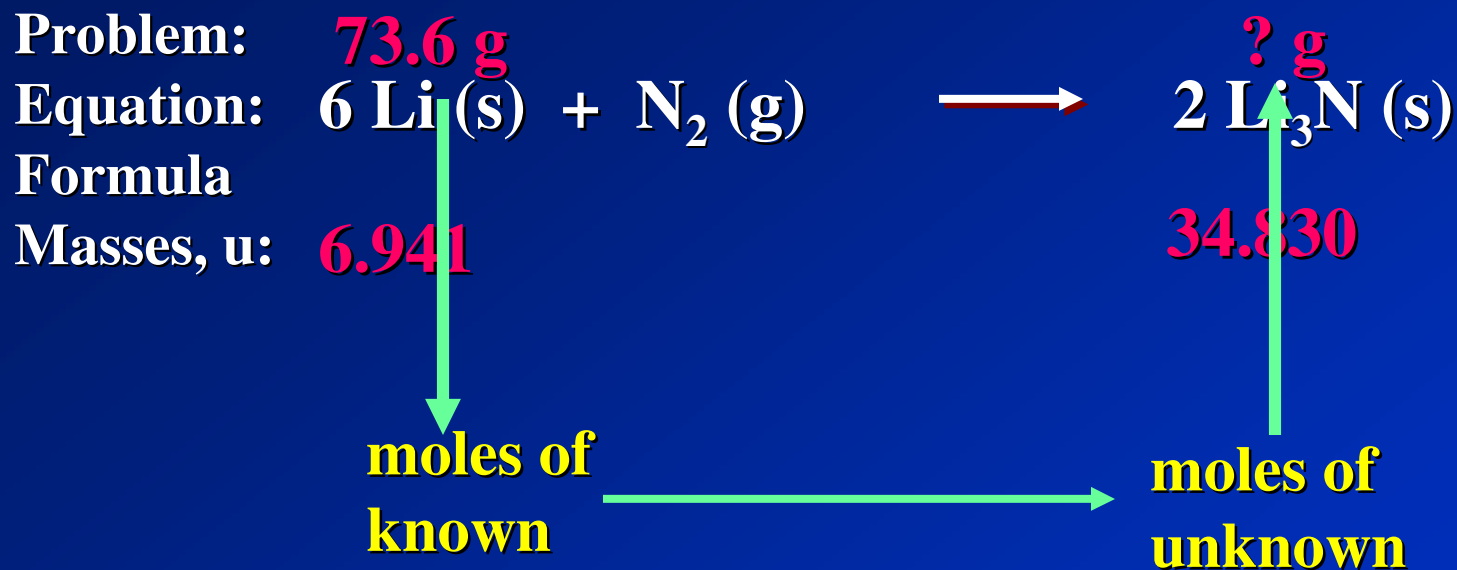
moles of
known

moles of
unknown

$$\frac{(58.63 \text{ g } \text{C}_3\text{H}_8) \cancel{(1 \text{ mol } \text{C}_3\text{H}_8)} \cancel{(4 \text{ mol } \text{H}_2\text{O})} (18.0153 \text{ g } \text{H}_2\text{O})}{\cancel{(44.097 \text{ g } \text{C}_3\text{H}_8)} \cancel{(1 \text{ mol } \text{C}_3\text{H}_8)} \cancel{(1 \text{ mol } \text{H}_2\text{O})}}$$

$$= 95.81 \text{ g } \text{H}_2\text{O}$$

How many grams of Li_3N (s) are formed when 73.6 g of Li (s) react with an excess of N_2 (g)?

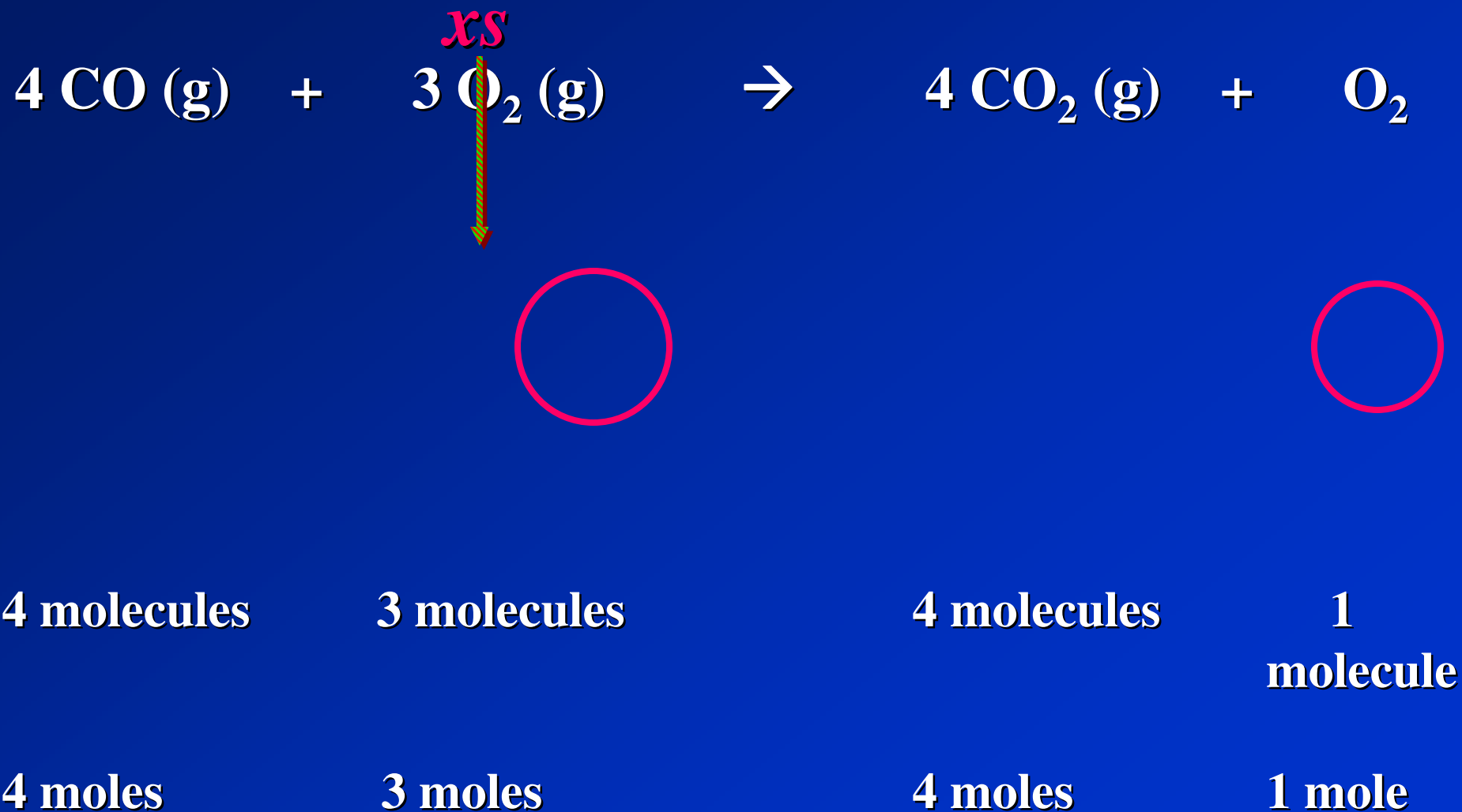


$$(73.6 \text{ g } \cancel{\text{Li}}) \frac{(1 \text{ mol } \cancel{\text{Li}})}{(6.941 \text{ g } \cancel{\text{Li}})} \frac{(2 \text{ mol } \cancel{\text{Li}_3\text{N}})}{(6 \text{ mol } \cancel{\text{Li}})} \frac{(34.830 \text{ g } \text{Li}_3\text{N})}{(\text{mol } \cancel{\text{Li}_3\text{N}})}$$

$$= \boxed{123 \text{ g } \text{Li}_3\text{N}}$$

The "Meaning" of a Chemical Equation

Limiting Reactant Problems



Limiting Reactant (*Limiting Reagent*): The reactant in a chemical reaction that is completely used up.

Excess Reactant (*Excess Reagent*): The reactant that is left over following a chemical reaction.



**Limiting
Reactant**

0.00 g
left over

**Excess
Reactant**

21.2 g
left over

When the hydrogen is all used up, the reaction HAS TO stop since no more of that reactant is available to react.

Limiting Reactant Problems

Making Pizza!

Due to your enormous personal popularity, you have just been elected *Social Chairperson* of the class and are required to prepare hot dog meals for the class. You find, however, upon going to the grocery store that you have just enough money to buy **5 packages** of hot dogs and **4 packages** of buns. Each package of hot dogs contains **12 hot dogs** but each package of buns contains **18 buns**. If **62 people** are in the class, will you have enough meals for everyone?

Let 1 meal = 1 hot dog (Hd) + 1 Bun (Bu)

Problem:	5 pkg	4 pkg	? meal
Equation:	Hd (s)	+ Bu (s)	→ Meal
Items/pkg:	12	18	

1. Find the reactant that is limiting:

$$(\cancel{5 \text{ pkg Hd}})(\cancel{12 \text{ Hd/pkg Hd}}) = \mathbf{60 \text{ Hd}}$$

$$(4 \text{ pkg Bu})(18 \text{ Bu/pkg Bu}) = 72 \text{ Bu}$$

Assume that Bu is the limiting reactant -

$$(72 \text{ Bu})(1 \text{ Hd/1 Bu}) = 72 \text{ Hd REQUIRED!!}$$

But Wait!! Only 60 Hd are AVAILABLE!! Thus, Hd, is, in fact, the limiting reactant.

2. Use the limiting reactant to find the amount of product that will be produced:

$$(60 \text{ Hd})(1 \text{ meal/1 Hd}) = 60 \text{ meals}$$

Since $60 < 62$, there will NOT be enough meals to go around and your popularity will fall off.

3. If asked, calculate the amount of excess reactant left over:

$$(60 \text{ Hd})(1 \text{ Bu}/1 \text{ Hd}) = 60 \text{ Bu REQUIRED!}$$

However, 72 buns are actually available -

$$72 \text{ Bu available} - 60 \text{ Bu required} = \mathbf{12 \text{ Bu in excess!}}$$

Thus, 12 will have to be wasted and the Faculty Advisor will be upset with you for wasting money.

Perhaps your days as the *Social Chairperson* are limited!!

Practice Exercise: A 100.00 g sample of Al (s) is heated with 100.00 g of O₂ (g). How many grams of Al₂O₃ (s) will form? How many grams of the excess reactant will be left over?

1. Set up the information table:

Problem:	100.00 g	100.00 g	? g
Equation:	4 Al (s)	+ 3 O ₂ (g)	→ 2 Al ₂ O ₃ (s)
Formula			
Masses, u:	26.9815	31.9988	101.9612

2. Find the limiting and excess reactants:

(“When in doubt, calculate moles”)

$$(100.00 \text{ g Al})(1 \text{ mol Al}/26.9815 \text{ g Al}) = 3.706 \underline{24} \text{ mol Al}$$

$$(100.00 \text{ g O}_2)(1 \text{ mol O}_2/31.9988 \text{ g O}_2) = 3.125 \underline{12} \text{ mol O}_2$$

Assume that O₂ is the limiting reactant:

$$(3.125 \text{ mol } \cancel{\text{O}_2})(4 \text{ mol Al}/3 \cancel{\text{ mol O}_2}) = 4.166 \text{ mol Al NEEDED}$$

This is more than the Al (s) AVAILABLE! Therefore, Al is the limiting reactant and O₂ is in excess.

3. Use the limiting reactant (Al) to calculate the grams of Al₂O₃ (s) that will be formed:

$$(3.706 \text{ mol } \cancel{\text{Al}})(2 \text{ mol Al}_2\text{O}_3/4 \cancel{\text{ mol Al}})(101.9612 \text{ g Al}_2\text{O}_3/\cancel{\text{ mol Al}_2\text{O}_3})$$
$$= \boxed{188.94 \text{ g Al}_2\text{O}_3}$$

4. Calculate the grams of excess reactant left over:

$$(3.706 \text{ mol Al})(3 \text{ mol O}_2 / 4 \text{ mol Al})(31.9988 \text{ g O}_2 / \text{mol O}_2)$$

$$= 88.9464 \text{ g O}_2 \text{ consumed}$$

$$100.00 \text{ g O}_2 \text{ available} - 88.9464 \text{ g O}_2 \text{ consumed}$$

$$= 11.05 \text{ g O}_2 \text{ excess}$$

Calculating Yields

Theoretical Yield - The amount of a product of a reaction calculated from the balanced chemical equation.

Actual Yield - The amount of a product actually produced in a chemical reaction.

Percentage Yield -

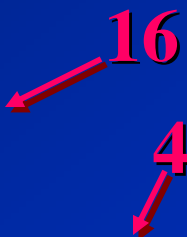
$$\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

20

Theoretical Yield = 20 popped

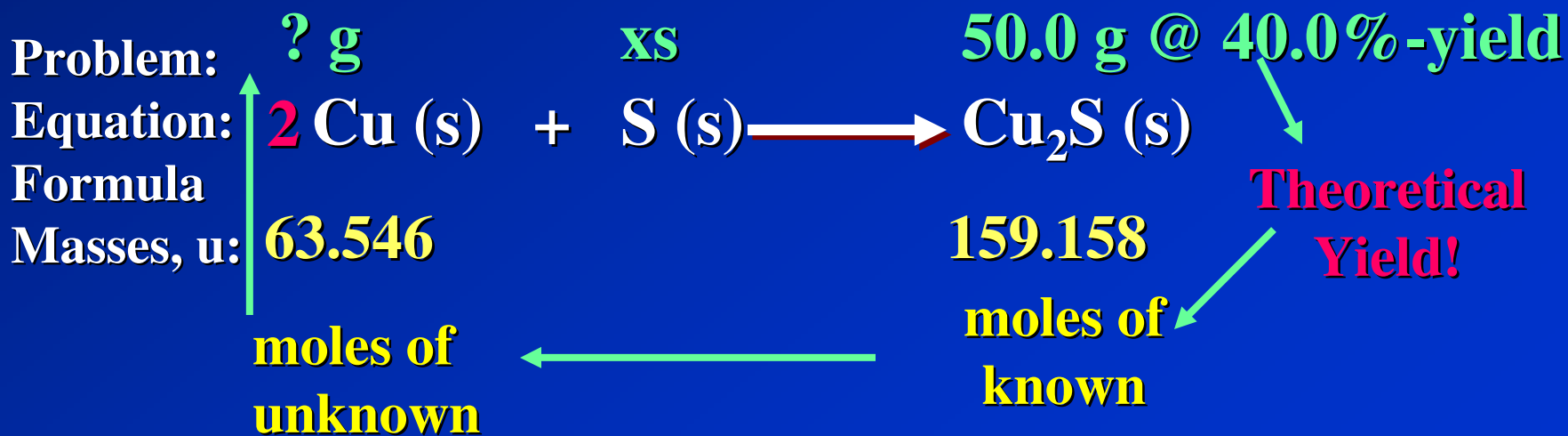
Actual Yield = 16 popped

%-Yield = (16/20)(100%) = 80%



Let's Practice!

Practice Exercise: Copper metal is heated with an excess of sulfur to produce copper(I) sulfide. If the % -yield is 40.0%, how many grams of copper are required to produce an actual yield of 50.0 g of Cu_2S ?



“When in doubt, calculate moles”

$$AY = 50.0 \text{ g} \quad \% - Y = 40.0\% \quad TY = ?$$

$$\frac{AY}{TY} \times 100\% = \% - Y \quad \Bigg| \quad TY = \frac{AY \times 100\%}{\% - Y}$$

$$TY = \frac{(50.0 \text{ g})(100\%)}{(40.0\%)} = 125.0 \text{ g Cu}_2\text{S}$$

$$(\cancel{125.0 \text{ g Cu}_2\text{S}})(1 \text{ mol Cu}_2\text{S}/\cancel{159.158 \text{ g Cu}_2\text{S}}) = 0.7854 \text{ mol Cu}_2\text{S}$$

$$(\cancel{0.7854 \text{ mol Cu}_2\text{S}})(\cancel{2 \text{ mol Cu}/1 \text{ mol Cu}_2\text{S}})(\cancel{63.546 \text{ g Cu/mol Cu}})$$

$$= \boxed{99.8 \text{ g Cu}}$$

Aqueous-Based Chemistry

Memorization!

Table 4.1

Acids and Bases

Memorization!

Table 4.2

Displacement Reactions

Reaction Types

Formation
(Redox)

Decomposition
*(Sometimes
Redox)*

Single Displacement
Single Replacement
(Redox)

Other

Metathesis
Double Displacement
Double Replacement
(Never Redox)

Important Definitions

Solution - A homogeneous mixture consisting of one or more substances uniformly dispersed as separate atoms, molecules, or ions in another substance.

Solvent - The component of a solution that is the “dissolving” medium. The solvent determines the physical state of the solution (*solid, liquid, or gas*).

Solute - The components of a solution that are “dissolved” by the medium.

Aqueous Solution - A solution wherein water is the solvent.

When the solute and solvent are both in the same physical state, the one in the largest quantity is the solvent.

Solubility - The amount of *solute* that will dissolve in a given quantity of *solvent* at a given temperature.

Saturated Solution - A solution that contains an amount of *solute* that is equal to its solubility.

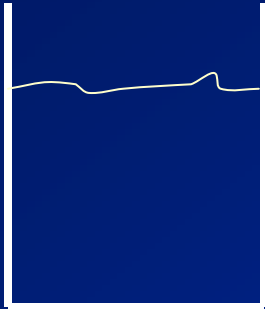


*The dissolved solute in the solution is in **dynamic equilibrium** with the undissolved solute (the precipitate).*

Dissolved solute \rightleftharpoons Precipitate

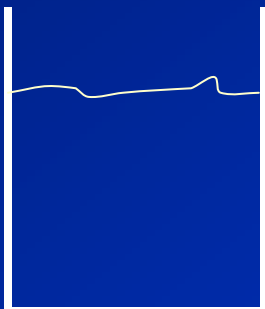
*The rate of dissolving is equal to the rate of precipitation.
This is a **dynamic equilibrium**.*

Unsaturated Solution - A solution that contains an amount of ***solute*** that is less than its solubility.



*All of the **solute** present is dissolved in an unsaturated solution.*

Supersaturated Solution - A solution that contains an amount of ***solute*** that is more than its solubility. This is a “metastable” state (“without stability”).



*If disturbed in anyway, the excess solute will precipitate out of solution and a **saturated solution** will result.*

Water is POLAR!

O

H

H

Practice Problem: If 50 mL of water at 10°C contain 0.000 0445 g of dissolved AgCl, (a) is the solution unsaturated, saturated, or supersaturated? (b) is the solution “dilute” or “concentrated”?

The solubility of AgCl in water at 10°C is 0.000 089 g AgCl/100 mL H₂O.

$$\begin{aligned} & (0.000\ 0445\ \text{g AgCl}/50\ \text{mL H}_2\text{O})(100\ \text{mL H}_2\text{O}/100\ \text{mL H}_2\text{O}) \\ & = 0.000\ 089\ \text{g AgCl}/100\ \text{mL H}_2\text{O} \end{aligned}$$

(a) Saturated

(b) Dilute

Solutes in Aqueous Solutions

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graph TD; A[Solutes in Aqueous Solutions] --- B[Non-Electrolytes]; A --- C[Electrolytes]
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Non-Electrolytes

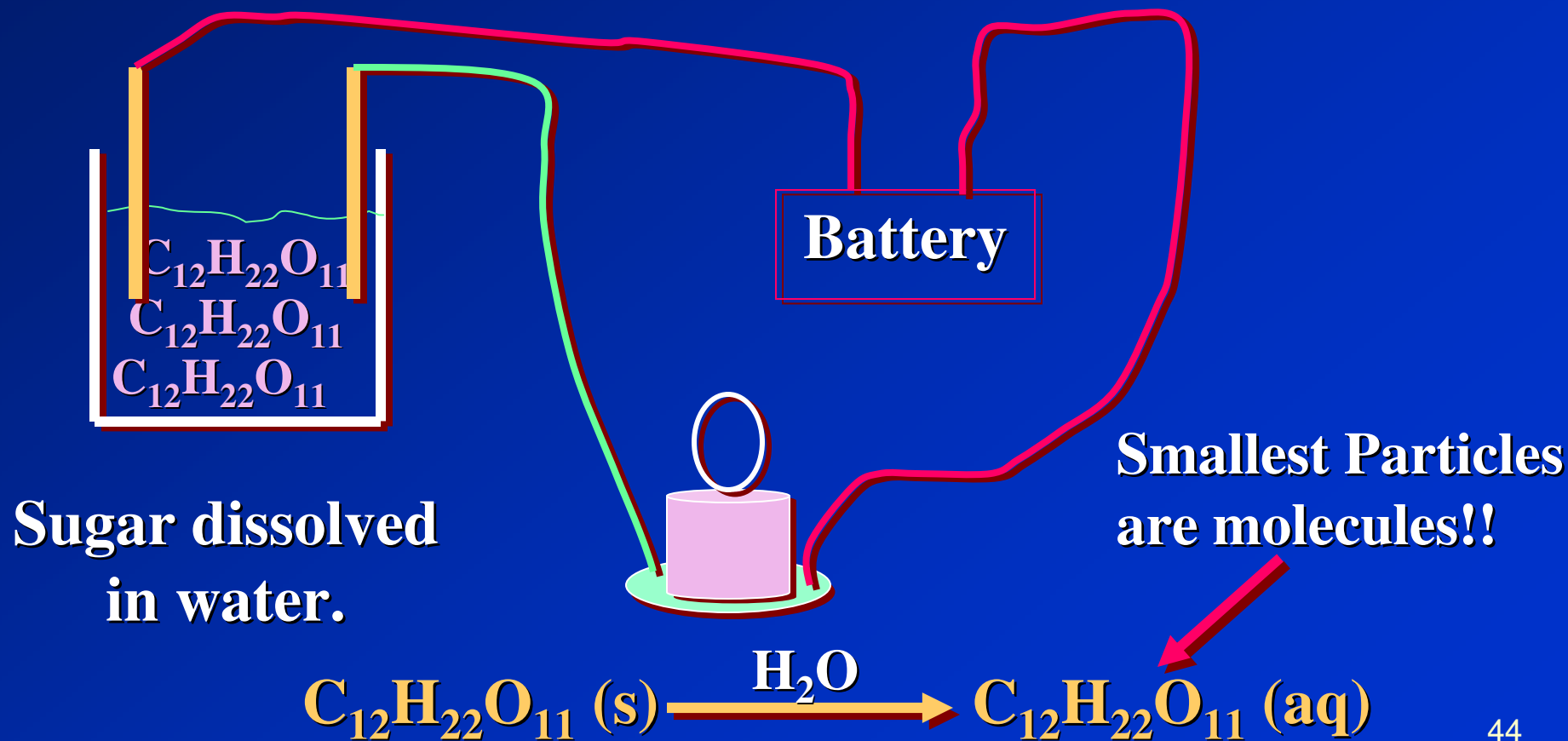
Solutes that do **NOT** produce ions when dissolved.

Electrolytes

Solutes that **DO** produce ions when dissolved.

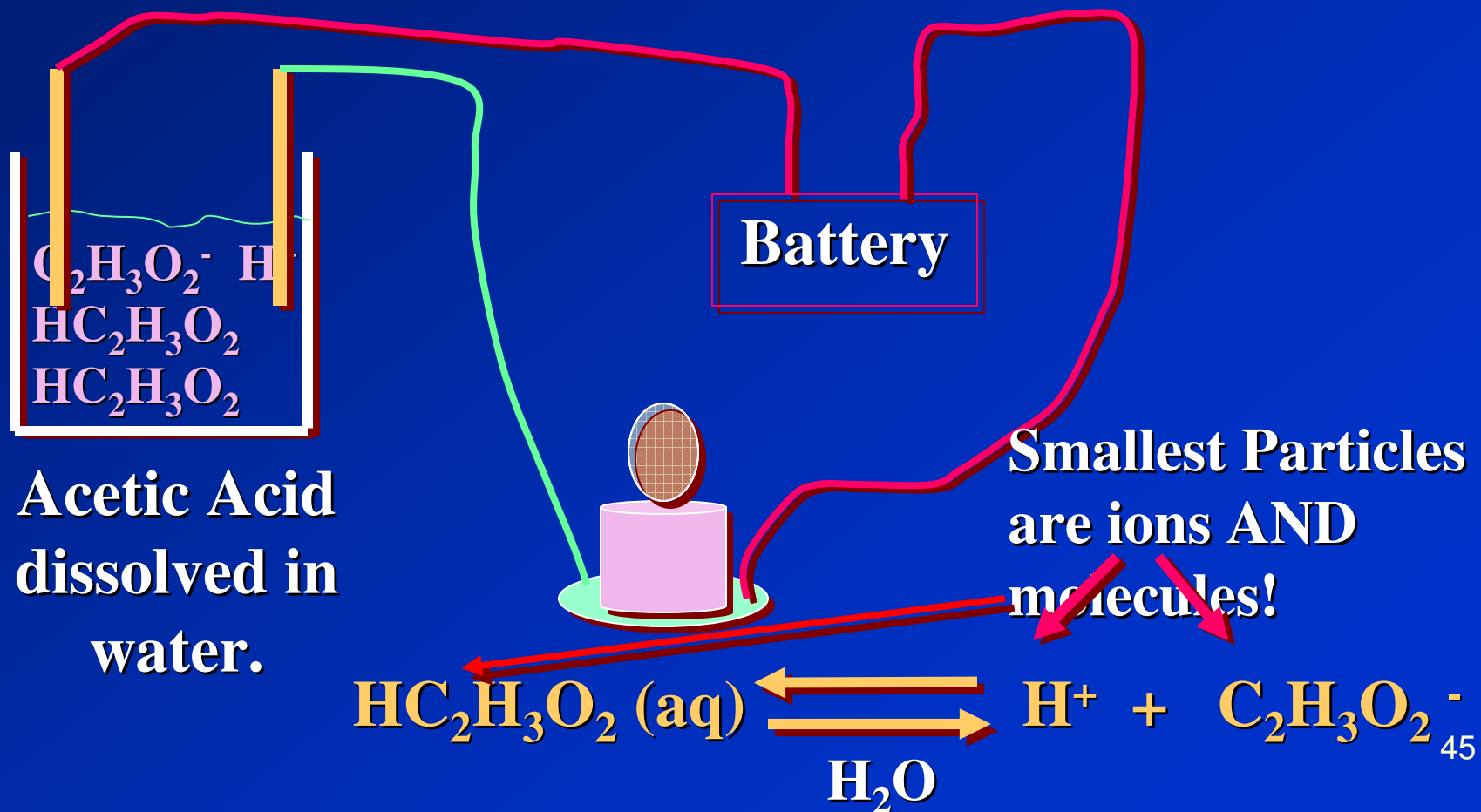
Properties of Solutes in Aqueous Solutions

Non-Electrolyte - A compound that does NOT produce ions when melted or dissolved in water. An aqueous solution of a non-electrolyte will NOT conduct electricity.



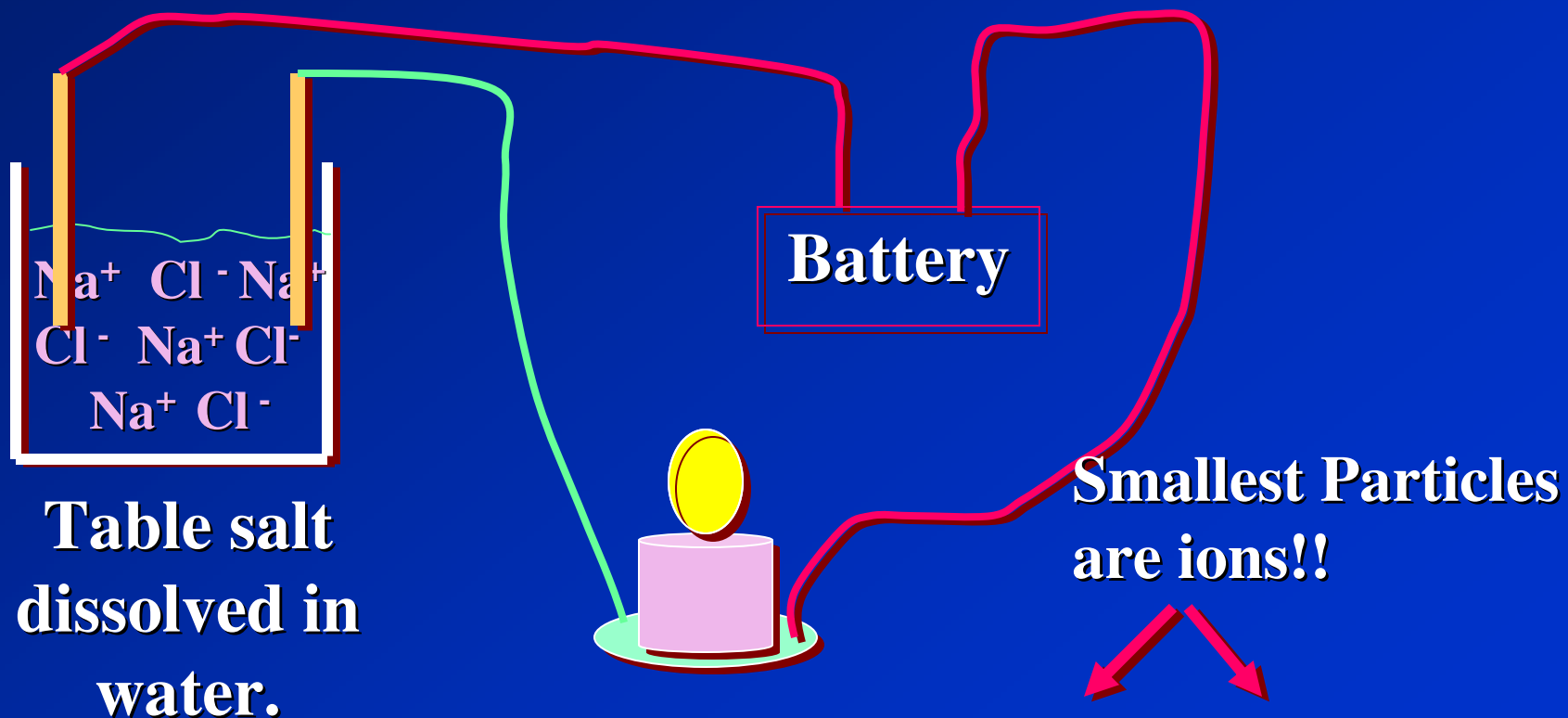
Properties of Solutes in Aqueous Solutions

Weak Electrolyte - A compound that produces only a few ions when dissolved in water. An aqueous solution of a weak electrolyte will conduct *some* electricity through the movement of the ions in the solution.



Properties of Solutes in Aqueous Solutions

Electrolyte - A compound that produces ions when melted or dissolved in water. An aqueous solution of an electrolyte will conduct electricity through the movement of the ions in the solution.



Electrolytes



Strong Electrolyte - An electrolyte that is completely or almost completely ionized (dissociated) in solution.



Weak Electrolyte - An electrolyte that ionizes in water only to a limited extent.



Acid - Acids are substances that ionize in H₂O to form H₃O⁺ - usually abbreviated simply as “H⁺”.

(This concept was developed by Arrhenius)



Strong Acid



Weak Acid

Or more simply...



http://nobelprize.org/nobel_prizes/chemistry/laureates/1903/arrhenius-lecture.html

Strong Acids: *Acids that dissociate completely into ions when placed in water.*

Strong Bases: *Bases that completely dissociate into ions to form ions in aqueous solution.*



$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

What is the pH of a 0.10 M solution of HCl? What about that of a 0.10 M solution of NaOH?

Since these are both strong electrolytes, the concentration of ions can be considered to be the same as the concentration of the solute.

Therefore: $[H_3O^+] = 0.10\text{ M}$ in the acid.

$$\text{pH} = -\log [H_3O^+] = -\log (0.10\text{ M}) = 1.00$$

$[OH^-] = 0.10\text{ M}$ in the base.

$$\text{pOH} = -\log [OH^-] = -\log (0.10\text{ M}) = 1.00$$

STRONG ACID

WEAK ACID

$$\text{pOH} + \text{pH} = 14.00$$

Therefore: $\text{pH} = 14.00 - \text{pOH} = 14.00 - 1.00$

$$= \boxed{13.00}$$

Weak Acids: *Acids that dissociate only to a limited extent into ions when placed in water.*

Weak Bases: *Bases that dissociate or react with water only to a limited extent into ions when placed in water.*

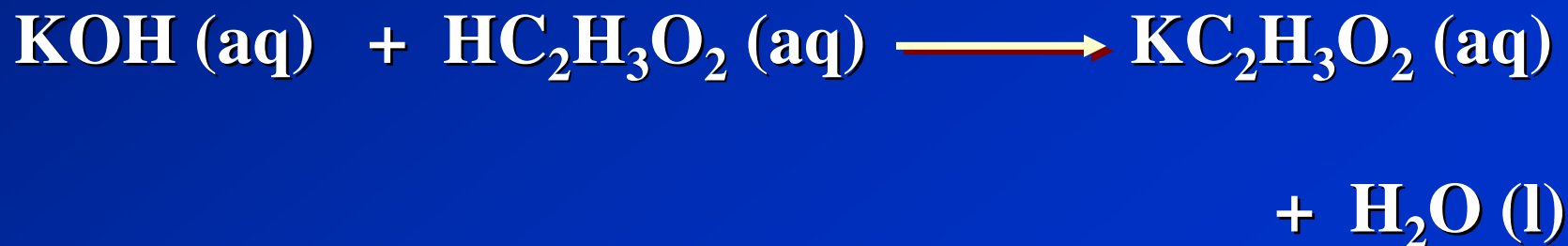
Base - Bases are substances that produce OH⁻ in aqueous solution.

(According to Mr. Arrhenius)



Salt - An ionic compound containing neither H⁺ ion nor OH⁻ ion. It is formed by the reaction of an acid and a base in a reaction known as “acid-base neutralization”.

Acid-Base Neutralization: The reaction between an acid and a base to produce water and a *Salt*.



State whether the following compounds are acids, bases, or salts:

KNO_3 - Salt

$\text{Cr}(\text{OH})_3$ - Base

NaOH - Base

H_2SO_4 - Acid

CaCl_2 - Salt

NH_3 - Base

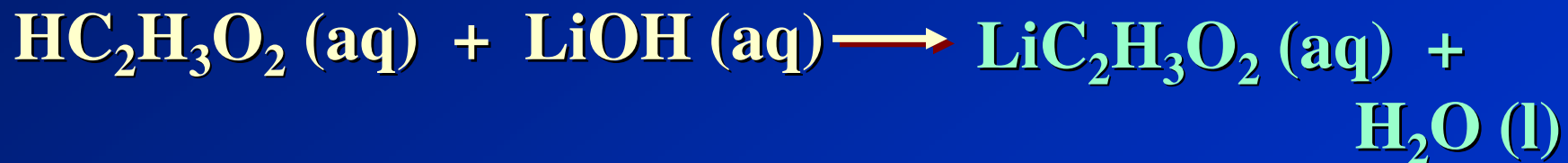
$\text{Ba}_3(\text{PO}_4)_2$ - Salt

Na_2SO_4 - Salt

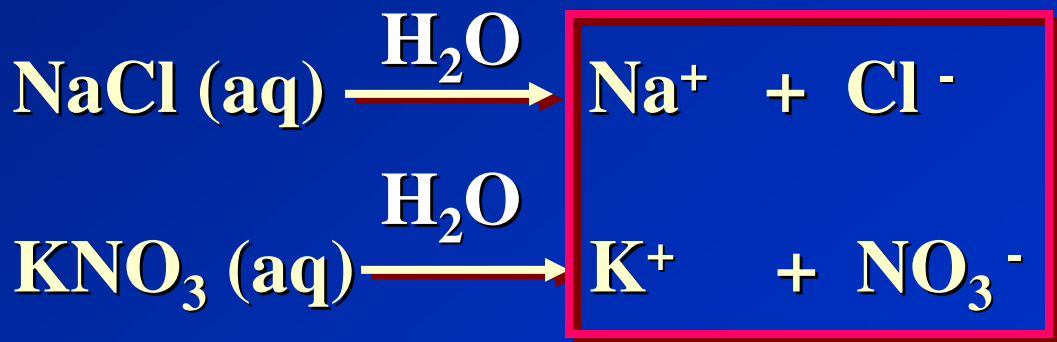
HCl - Acid

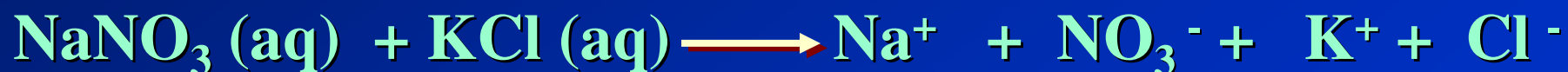
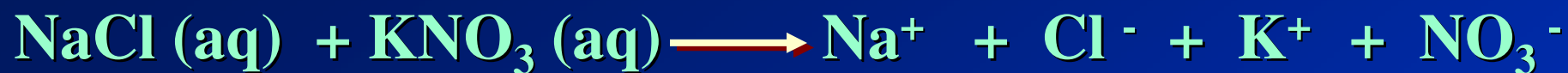
**H_2O - Acid and
Base**

Complete the following neutralization equations:



Ionic Solutions - Solutions containing ions dispersed in some solvent.





Even though we started with two different salts, the two solutions have exactly the same composition. The cations and anions of the two solutions are identical.

Ion Attractions – Who's Attracted to Whom?

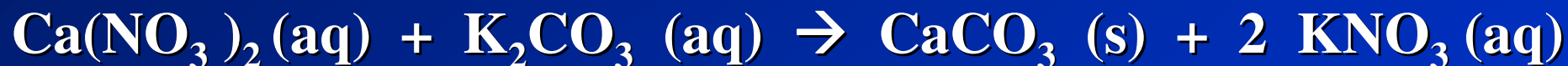
Precipitation Reactions



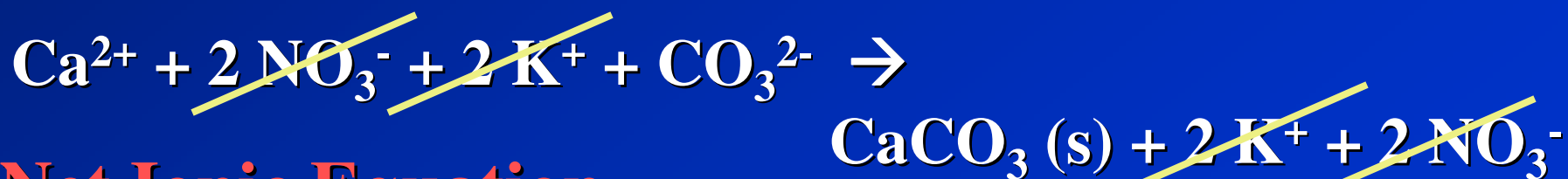
Precipitation Reactions

- Form an insoluble compound

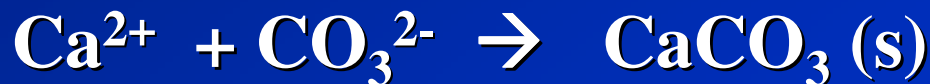
Molecular Equation



Complete Ionic Reaction

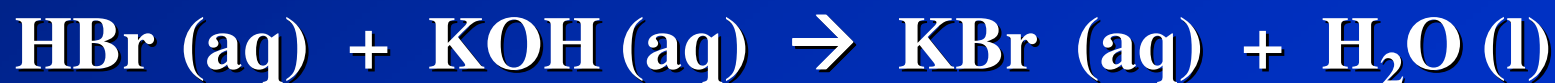


Net Ionic Equation

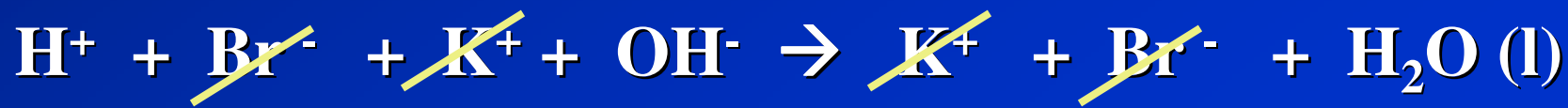


Acid-Base Reactions

- Acid + Base \longrightarrow Salt + Water
- Recall the acids and bases from previous discussion
- **“Molecular” equation**

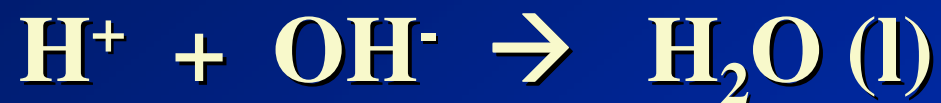


- **Complete ionic equation**



Acid-Base Neutralization

□ **Net Ionic Equation**

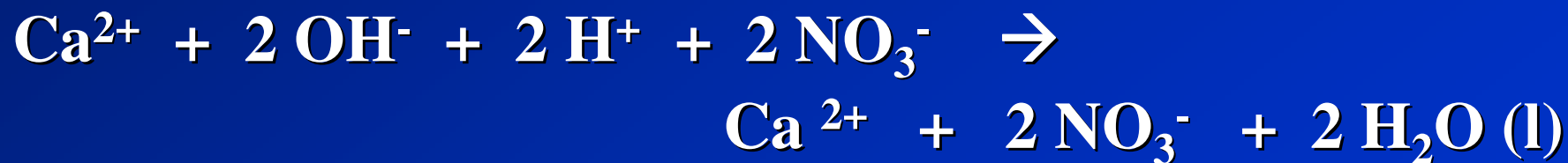


Acid-Base Neutralization

□ “Molecular” Equation



□ Complete Ionic Equation



□ Net Ionic Equation



or



Displacement Reactions

Reaction Types

Formation
Combination
(Redox)

Decomposition
*(Sometimes
Redox)*

Single Displacement
Single Replacement
(Redox)

Other

Metathesis
Double Displacement
Double Replacement
(Never Redox)

Displacement Reactions

Single Displacement

Occurs when one element is more **ACTIVE** than another

Double Displacement

Occurs when ionic compounds switch partners and forms a

↳ **precipitate**

↳ **gas** or

↳ **weak or non-electrolyte**

Displacement Reactions (*Single Replacement*)

How would you know that Cu is more **ACTIVE** than Ag?

Look at the

ACTIVITY SERIES

K
Na
Mg
Al
Zn
Cr
Fe
Cd
Co
Ni
Sn
Pb
H
Cu
Hg
Ag
Pt
Au

Displacement Reactions (*Single Replacement*)



“Molecular Equation”



Complete (total) Ionic Equation



Net Ionic Equation



Will this reaction actually occur??

K
Na
Mg
Al
Zn
Cr
Fe
Cd
Co
Ni
Sn
Pb
H
Cu
Hg
Ag
Pt
Au

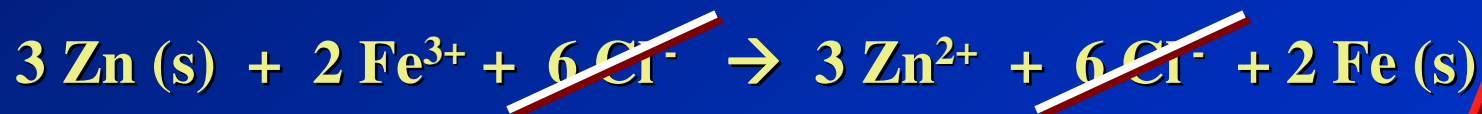
Displacement Reactions (*Single Replacement*)



“Molecular Equation”



Complete (total) Ionic Equation



Net Ionic Equation



Will this reaction actually occur??

K
Na
Mg
Al
Zn
Cr
Fe
Cd
Co
Ni
Sn
Pb
H
Cu
Hg
Ag
Pt
Au

Displacement Reactions (*Single Replacement*)



“Molecular Equation”



Complete (total) Ionic Equation

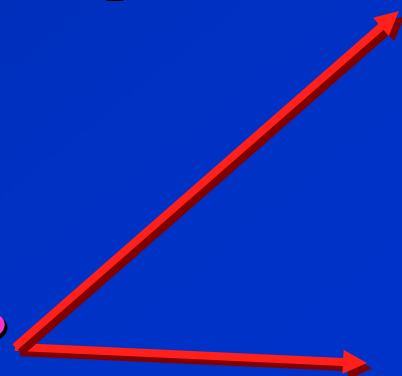


Net Ionic Equation



Will this reaction actually occur??

K
Na
Mg
Al
Zn
Cr
Fe
Cd
Co
Ni
Sn
Pb
H
Cu
Hg
Ag
Pt
Au
69



Expressing Concentration

Concentration - The amount of **solute** present in a given quantity of **solvent** or **solution**.

1.
$$\frac{\text{mass solute}}{\text{volume solvent}} \quad \frac{35.7 \text{ g NaCl}}{100 \text{ mL H}_2\text{O}}$$

(Typical way to express solubilities)

2. Percent by Mass

$$\left[\frac{\text{mass solute}}{\text{mass solution}} \right] (100 \%)$$

What is the percent by mass of sugar in a solution containing 30.0 g of sugar and 65.0 g H₂O?

(mass of solute/mass of solution)(100%) =

$$\left[\frac{(30.0 \text{ g sugar})}{(30.0 \text{ g sugar} + 65.0 \text{ g H}_2)} \right] (100.0\%) =$$

31.6 % by mass

3. Molarity (M)

$$\text{Molarity} = \frac{\text{moles solute}}{\text{liter solution}}$$

Practice Problem: A 750-mL solution contains 1.22 mol of KNO_3 . Calculate the molarity of the solution.

$$\left[\frac{1.22 \text{ mol KNO}_3}{750 \text{ mL soln}} \right] \left[\frac{1000 \text{ mL soln}}{\text{L soln}} \right] = 1.63 \text{ M KNO}_3$$

Practice Problem: Calculate the molarity of a solution that contains 5.25 g AgNO₃ in 125 mL of solution.

1. Calculate the formula mass of AgNO₃.

$$1 \text{ Ag@ } 107.868 \text{ u} = 107.868 \text{ u}$$

$$1 \text{ N@ } 14.0067 \text{ u} = 14.0067 \text{ u}$$

$$3 \text{ O@ } 15.9994 \text{ u} = 47.9982 \text{ u}$$

$$1 \text{ AgNO}_3 = 169.873 \text{ u}$$

2. Determine the moles of AgNO₃ in 5.25 g AgNO₃.

$$(5.25 \text{ g AgNO}_3) \left(\frac{1 \text{ mol AgNO}_3}{169.873 \text{ g AgNO}_3} \right) =$$

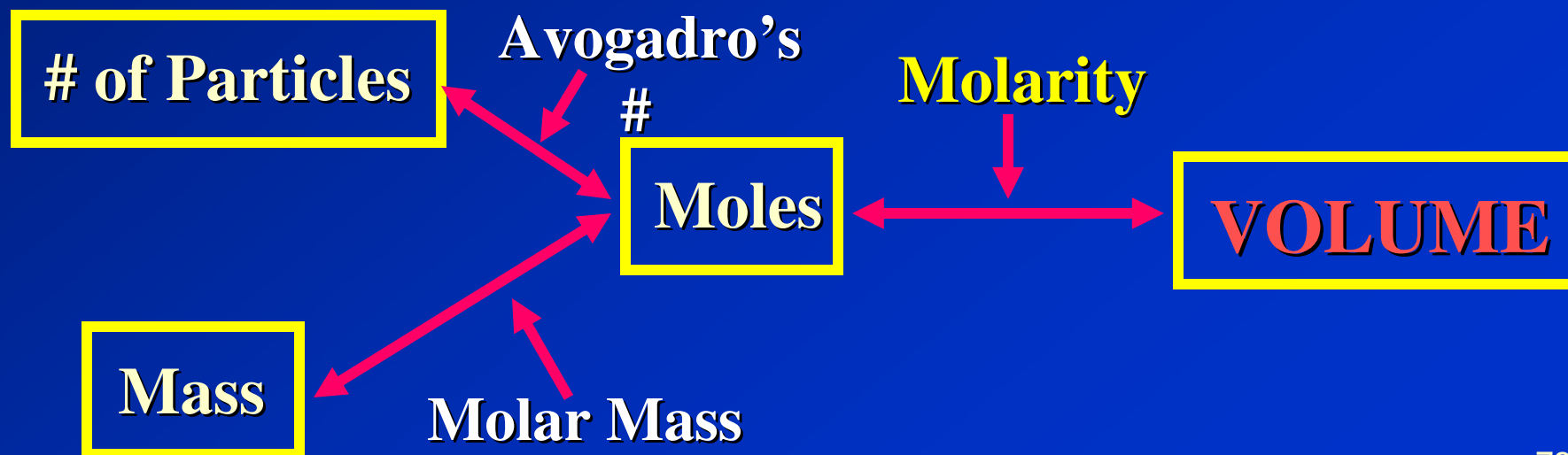
$$0.0309 \text{ mol AgNO}_3$$

3. From the moles of AgNO_3 and the mL of solution, calculate the molarity of the solution.

$$\left[\frac{0.03091 \text{ mol AgNO}_3}{125 \text{ mL soln}} \right] \left[\frac{1000 \text{ mL soln}}{1 \text{ L soln}} \right]$$

$$= 0.247 \text{ M AgNO}_3$$

(“When in doubt, calculate moles”)



Practice Problem: Calculate the molarity of chloride ions in a 0.125 M solution of AlCl_3 .



Every mole of AlCl_3 produces 3 moles of chloride ions.

$$\begin{aligned} & (\cancel{0.125 \text{ mol AlCl}_3/\text{L soln}})(3 \text{ mol Cl}^-/\cancel{\text{mol AlCl}_3}) \\ & = \mathbf{0.375 \text{ M Cl}^-} \end{aligned}$$

Practice Problem: How many liters of 0.20 M AlCl_3 will contain 0.90 mol Cl^- ?

$$(\cancel{0.90 \text{ mol Cl}^-})(1 \text{ mol AlCl}_3/3 \cancel{\text{ mol Cl}^-}) = 0.30 \text{ mol AlCl}_3$$

$$(\cancel{0.30 \text{ mol AlCl}_3})(1 \text{ L soln}/\cancel{0.20 \text{ mol AlCl}_3}) = \mathbf{1.5 \text{ L soln}}$$

Practice Problem: How many grams of NaNO_3 are needed to make 250.0 mL of 0.325 M NaNO_3 ?

$$(\cancel{250.0 \text{ mL soln}})(1 \text{ L soln}/\cancel{10^3 \text{ mL soln}}) = 0.250 \text{ L soln}$$

$$(\cancel{0.250 \text{ L soln}})(\cancel{0.325 \text{ mol NaNO}_3/\text{L soln}}) = 0.08125 \text{ mol NaNO}_3$$

$$1 \text{ Na@ } 22.98977 \text{ u} = 22.98977 \text{ u}$$

$$1 \text{ N@ } 14.0067 \text{ u} = 14.0067 \text{ u}$$

$$\underline{3 \text{ O@ } 15.9994 \text{ u} = 47.9982 \text{ u}}$$

$$1 \text{ NaNO}_3 = 84.9947 \text{ u}$$

$$(\cancel{0.08125 \text{ mol NaNO}_3})(\cancel{84.9947 \text{ g NaNO}_3/\text{mol NaNO}_3})$$

$$= 6.91 \text{ g NaNO}_3$$

How would you make this solution?

- 1. Carefully weigh out 6.91 g of NaNO_3 into a 250-mL volumetric flask.**
- 2. Add between 50 and 100 mL of distilled water to the flask and swirl gently until all of the NaNO_3 has dissolved.**
- 3. When all of the NaNO_3 is dissolved. Add more water until the level is just below the fiduciary mark on the long neck of the volumetric flask.**
- 4. Add additional distilled water drop-wise until the bottom of the meniscus is EXACTLY on the mark.**

You now have 250 mL of a 0.325 M NaNO_3 solution!

Dilution

Solutions with a specific concentration of solute can be made by diluting more concentrated solutions that contain the same solute. The initial concentration must be accurately known.

Practice Problem: How many mL of 5.00 M $\text{K}_2\text{Cr}_2\text{O}_7$ solution must be diluted in order to prepare 250-mL of a 0.100 M solution?

$$(\cancel{250 \text{ mL dil. soln}})(1 \text{ L dil. soln}/\cancel{10^3 \text{ mL dil. soln}}) = 0.250 \text{ L dil. soln}$$

$$(\cancel{0.250 \text{ L dil. soln}})(\cancel{0.100 \text{ mol K}_2\text{Cr}_2\text{O}_7/\text{L dil. soln}}) =$$

0.0250 mol $\text{K}_2\text{Cr}_2\text{O}_7$ needed!!

$$(\cancel{0.0250 \text{ mol K}_2\text{Cr}_2\text{O}_7})(1 \text{ L conc. soln}/\cancel{5.00 \text{ mol K}_2\text{Cr}_2\text{O}_7})$$

$$= 0.00500 \text{ L conc. soln.}$$

$$(\cancel{0.00500 \text{ L conc. soln.}})(\cancel{10^3 \text{ mL conc. soln}/\cancel{\text{L conc. soln}}})$$

$$= \mathbf{5.00 \text{ mL conc. soln}}$$

Summary: Recognize that the moles of solute present in the sample of the concentrated solution are equal to the moles present after dilution. Dilution does NOT change that! Also recognize that **$M \times L = \text{moles}$** .

$$\therefore M_{\text{initial}} \times V_{\text{initial}} = M_{\text{final}} \times V_{\text{final}}$$

$$V_{\text{initial}} = (M_{\text{final}} \times V_{\text{final}})/M_{\text{initial}}$$

$$V_{\text{initial}} = \frac{(0.100 \text{ M} \times 0.250 \text{ L})(10^3 \text{ mL})}{(5.00 \text{ M}) (\text{ L})} = 5.00 \text{ mL}$$

How would you make this solution?

- 1. Carefully transfer exactly 5.00 mL of the concentrated solution to a 250-mL volumetric flask.**
- 2. Add between 50 and 100 mL of distilled water to the flask and swirl gently until the solution is homogeneous.**
- 3. Add more water until the level is just below the fiduciary mark on the long neck of the flask.**
- 4. Add additional distilled water drop-wise until the bottom of the meniscus is EXACTLY on the mark.**

*You now have 250 mL of a 0.100 M $K_2Cr_2O_7$ solution!*₇₉

Acid-Base Titration

Volumetric Analysis - An analytical procedure in which quantitative information is obtained by measuring the volume of reacting species.

Titration - A controlled addition of one solution to another.

Indicator - A dye that will change colors at the end point of a reaction.

End Point - The point in a titration when the indicator changes color signifying that the reaction is complete.

Standard Solution - A solution that has an accurately known concentration.

Acid-Base Titration

Practice Problem: A 20.15 mL sample of a HCl (aq) solution is titrated to a phenolphthalein end-point with 19.87 mL of a standard solution of 0.1246 M NaOH (aq). What is the concentration of the HCl (aq)?

20.15 mL	19.87 mL
?M	0.1246 M



19.87 mL NaOH	1 L NaOH	0.1246 mol NaOH
<hr/>	<hr/>	<hr/>
	10³ mL NaOH	L NaOH

$$= 2.475\text{8} \times 10^{-3} \text{ mol NaOH}$$

Acid-Base Titration

$2.4758 \times 10^{-3} \text{ mol NaOH}$	1 mol HCl	1
	1 mol NaOH	0.020 15 L HCl

= 0.1229 M HCl

Oxidation- Reduction - *Redox*

“It takes Two to Tango”

Oxidation-Reduction Reactions:

1. Definition - A reaction in which electrons are transferred from one substance to another; one substance is *oxidized* and the other substance is *reduced*.

2. Applications -

A. Batteries

B. Bleaches

C. Photography

D. Elements combining to form compounds

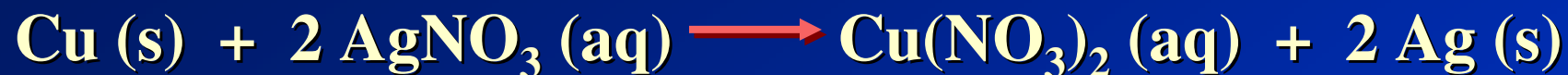
E. Elements replacing other elements in compounds

- 3. Oxidation** - A Loss of electrons.
- 4. Reduction** - A Gain of electrons; a “REDUCTION in positive charge”.
- 5. Oxidizing Agent** - The substance in a Redox reaction that is REDUCED! It Gains electrons.
- 6. Reducing Agent** - The substance in a Redox reaction that is OXIDIZED! It Loses electrons.
- 7. Examples** - $2 \text{H}_2 (\text{g}) + \text{O}_2 (\text{g}) \longrightarrow 2 \text{H}_2\text{O} (\text{l})$

**Reducing
Agent**

**Oxidizing
Agent**

7. Examples (*Continued*) -



Reducing Agent **Oxidizing Agent**



Reducing Agent **Oxidizing Agent**



Reducing Agent **Oxidizing Agent**

7. Examples (*Continued*) -



Reducing
Agent

Oxidizing
Agent



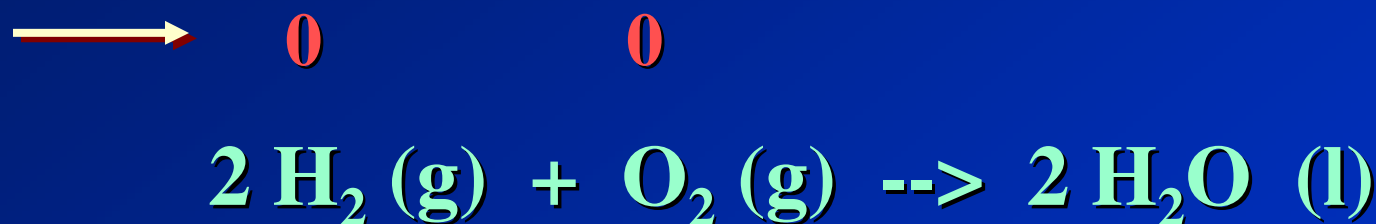
Reducing
Agent

Oxidizing
Agent

*Autooxidation or
Disproportionation
Reaction*

Oxidation Number Rules

1. The oxidation number of an atom in a free element (not in a compound!) is ALWAYS zero, regardless of its formula.

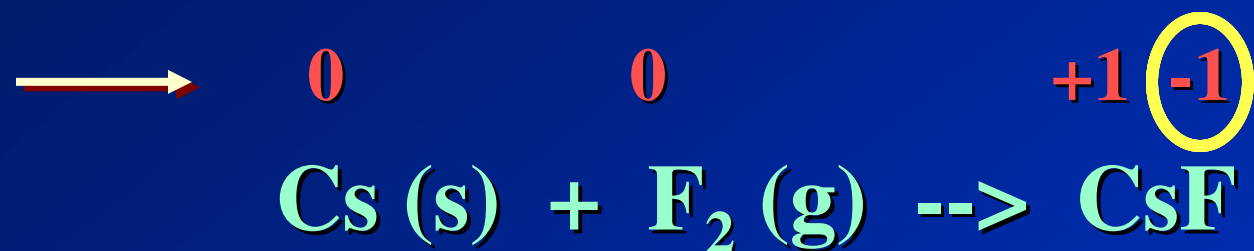


2. The oxidation number of a monatomic (simple) ion is equal to the charge on the ion.

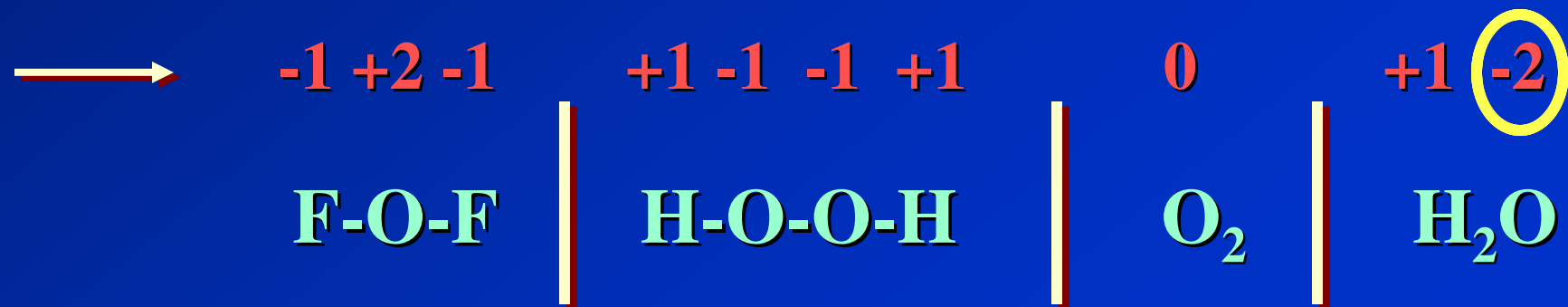


Oxidation Number Rules

3. The oxidation number of fluorine is ALWAYS **-1** in its compounds.

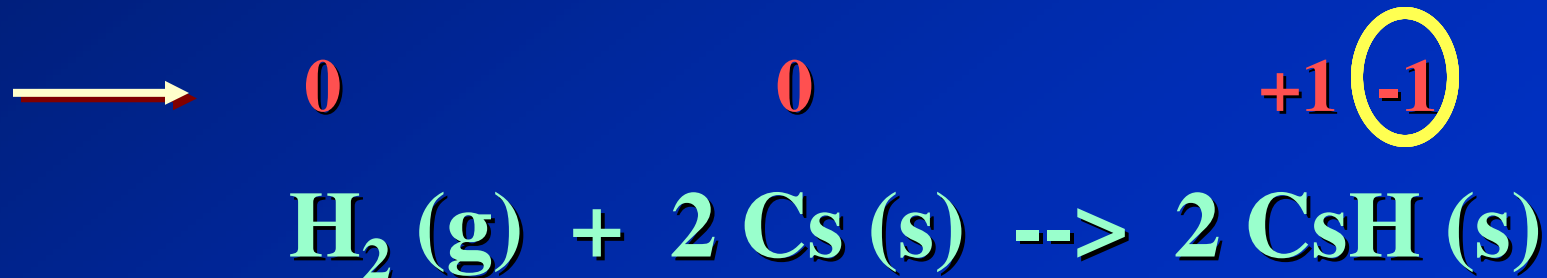
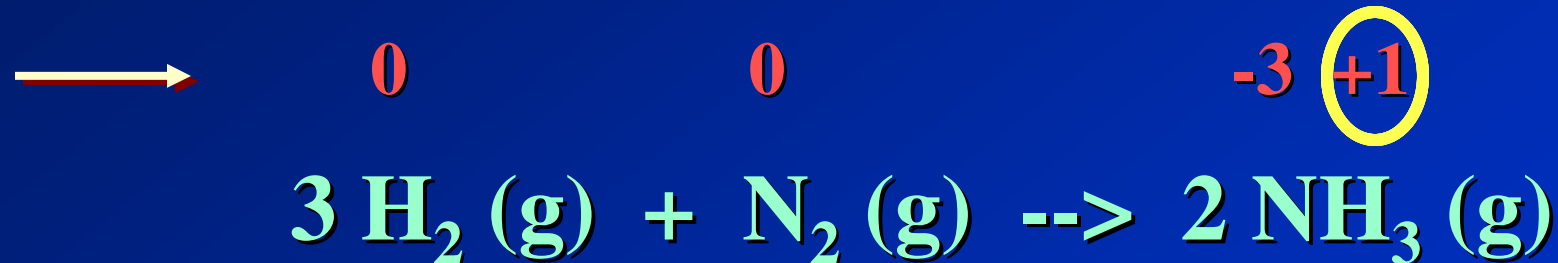


4. The oxidation number of oxygen is **-2** unless it is bonded to fluorine or itself.



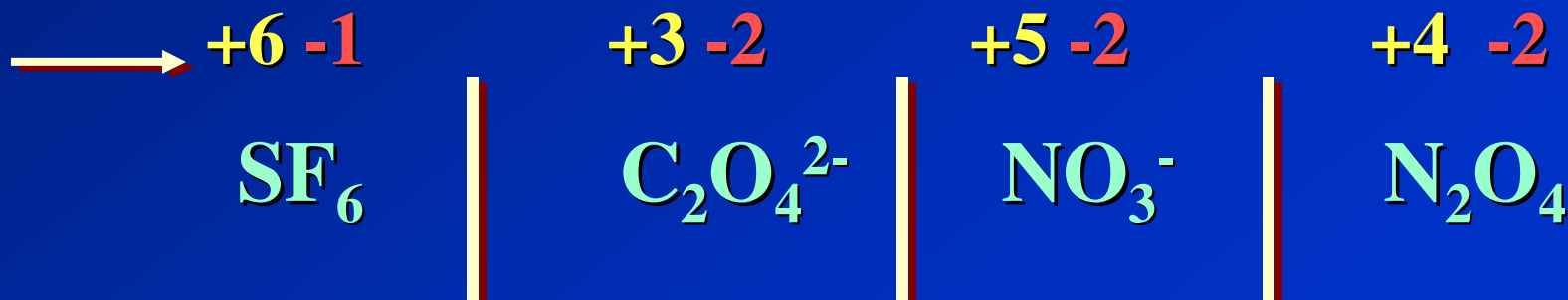
Oxidation Number Rules

5. Hydrogen has an oxidation state of **+1** in all compounds except in metal hydrides where it is **-1**.



Oxidation Number Rules

6. The sum of the oxidation numbers in a neutral compound is *zero*; the sum for a polyatomic ion is equal to the charge on the ion.



Oxidation Number Rules

7. Some Common Oxidation Numbers:

↖ Alkali Metals are ALWAYS +1

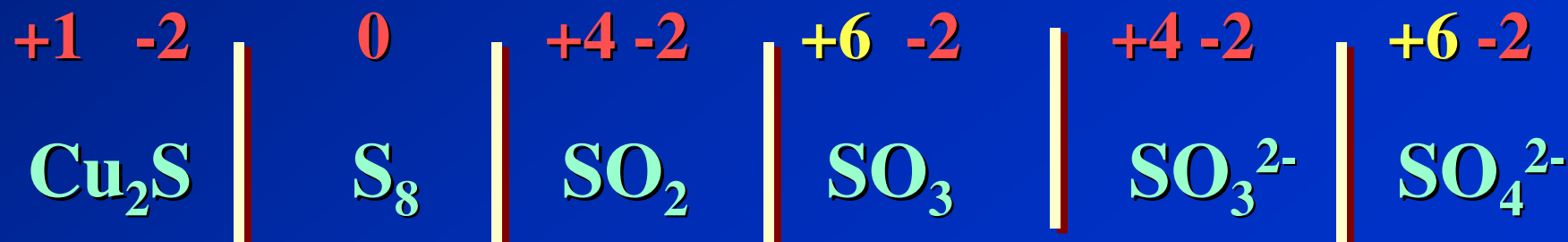
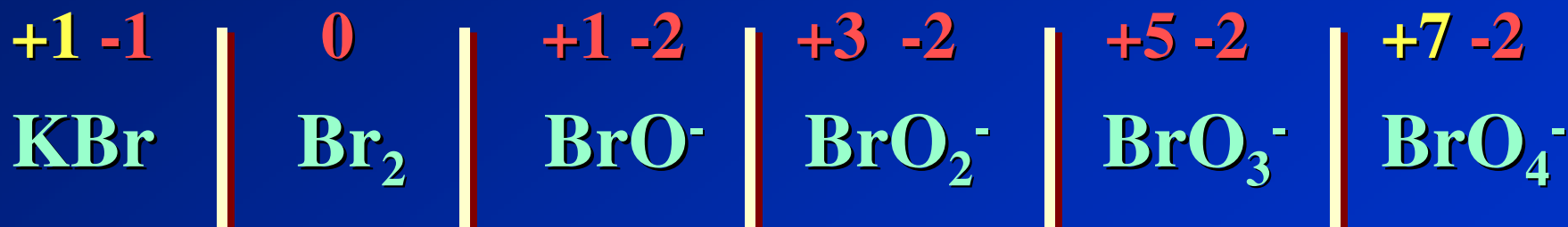
↖ Alkaline Earth Metals are ALWAYS +2

↖ Halogens are -1 except when bonded to oxygen or to a halogen above it in the Periodic Table



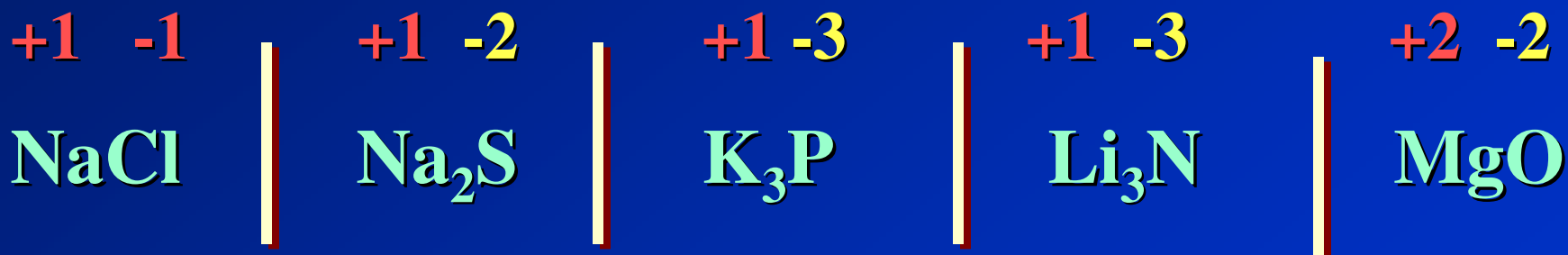
Oxidation Number Rules

8. For Main Group elements, the highest possible oxidation state is equal to its group number.



Oxidation Number Rules

9. For Main Group elements, the lowest possible oxidation state is equal to its group number minus 8.



Practice Exercises:



Oxidation Numbers

- Assign oxidation numbers to each element in these compounds and polyatomic ions:

