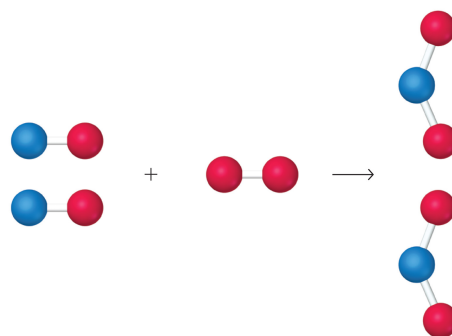
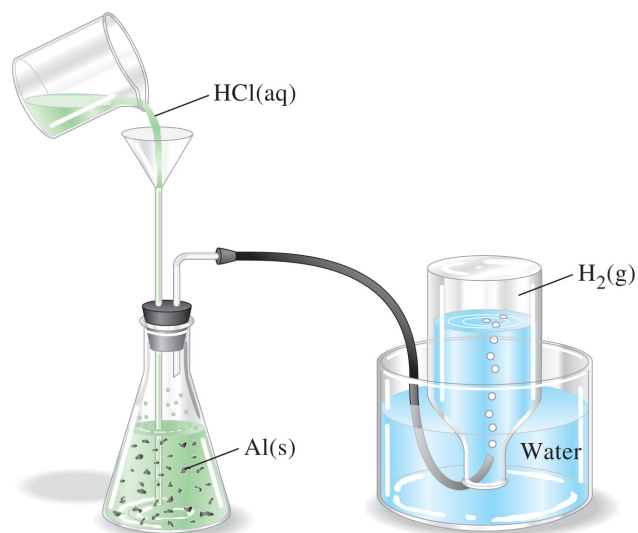


# Chapter 3: Stoichiometry of Formulae and Equations



# 3.1 Avogadro's number



Amedeo Avogadro  
(1776 – 1856)



The mole is an amount of substance that contains the same number of elementary entities as there are carbon-12 atoms in exactly 12 g of carbon-12.

$$N_A = 6.022\ 141\ 29 \times 10^{23} \text{ mol}^{-1}$$



# Molar Mass

- The molar mass, **M**, is the mass of one mole of a substance.

$$M \text{ (g/mol } ^{12}\text{C)} = A \text{ (g/atom } ^{12}\text{C)} \times N_A \text{ (atoms } ^{12}\text{C /mol } ^{12}\text{C)}$$

- the numerical value of the atomic mass and the molar mass of an element are identical, but they have different units!

atomic mass of Na = 22.99 \_\_\_\_\_  
molar mass of Na = 22.99 \_\_\_\_\_



# Molecular Mass from Atomic Masses

molecular mass = sum of atomic mass of each atom in the molecule (in amu)

What is the *molecular mass* of H<sub>2</sub>O?

$$2 \text{ atoms H} \times \frac{1.008 \text{ amu}}{1 \text{ H atom}} = 2.016 \text{ amu}$$

$$1 \text{ atom O} \times \frac{15.999 \text{ amu}}{1 \text{ O atom}} = 15.999 \text{ amu}$$

$$1 \text{ molecule H}_2\text{O} = 2.016 \text{ amu} + 15.999 \text{ amu} = 18.015 \text{ amu}$$



# Molar Mass vs. Molecular Mass

mass in grams of 1 mole of the molecule

What is the *molar mass* of H<sub>2</sub>O?

$$2 \text{ moles H} \times \frac{1.008 \text{ g}}{1 \text{ H mol}} = 2.016 \text{ g}$$

$$1 \text{ moles O} \times \frac{15.999 \text{ g}}{1 \text{ O mol}} = 15.999 \text{ g}$$

$$1 \text{ mole of H}_2\text{O} = 2.016 \text{ g} + 15.999 \text{ g} = 18.015 \text{ g}$$

**Molecular Mass:** mass of one molecule

18.015 **amu**

**Molar Mass:** mass of one mole of molecules

18.015 **g/mol**



# Example: Molar Masses

Potassium-40 is one of the few naturally occurring radioactive isotopes of elements of low atomic number. Its percent natural abundance among K isotopes is 0.012%. How many  $^{40}\text{K}$  atoms do you ingest by drinking one cup of whole milk containing 371 mg of K?

What we want: number of atoms of  $^{40}\text{K}$

What we have: natural abundance of  $^{40}\text{K}$ ,  
molar mass of K, the mass of K



# Percent Composition

- number of parts of a component in 100 parts of the whole
  - ex. 10% means “10 parts x per 100 parts of the whole”
- IMPORTANT: must be defined by a unit!
  - ex. a rock contains 3.5% gold **by mass** means 3.5 g of gold per 100 g of rock
  - ex. a bottle of wine contains 10.7% alcohol **by volume** means 10.7 mL of alcohol per 100 mL of wine



# Percent Composition

- when expressed as a conversion factor, the numerator and denominator must have the SAME UNITS
  - ex. a rock contains 3.5% gold **by mass**

$$\frac{3.5 \text{ g gold}}{100 \text{ g rock}} = \frac{3.5 \text{ kg gold}}{100 \text{ kg rock}} = \frac{3.5 \text{ oz gold}}{100 \text{ oz rock}}$$

- ex. a bottle of wine contains 10.7% alcohol **by volume**

$$\frac{10.7 \text{ mL EtOH}}{100 \text{ mL wine}} = \frac{10.7 \text{ L EtOH}}{100 \text{ L wine}} = \frac{10.7 \text{ tbsp EtOH}}{100 \text{ tbsp wine}}$$



# Your Turn...

Calculate the mass % of hydrogen in ammonium bicarbonate.

- A) 5.166 %
- B) 8.392 %
- C) 6.387 %
- D) 22.82 %
- E) I'm not sure.



# Example : Using percent composition

A solution of sucrose in water is 28.0% sucrose by mass and has a density of 1.118 g/mL. What mass of sucrose (in grams) is in 3.50 L of this solution?

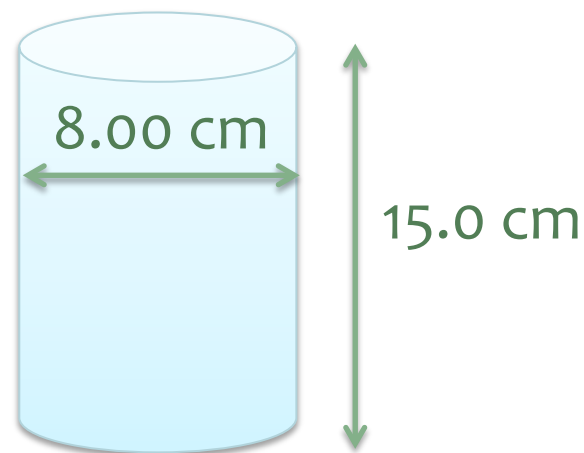


# Example : Integrative Problem

## 2010 Exam:

Barium sulfate is used as a radiocontrast agent for X-ray imaging of the digestive system. A patient is given a beaker filled with a solution that contains 0.25%  $\text{BaSO}_4$  by mass. If the density of the imaging solution is 1.45 g/mL, and given that the percent natural abundance of  $^{134}\text{Ba}$  is 2.417%, how many  $^{134}\text{Ba}$  atoms will be ingested by the patient?

*Recall:* volume of a cylinder =  $\pi r^2 h$



# 3.2 Chemical Formulas

## Empirical formula:

- simplest whole number RATIO of elements

## Molecular formula:

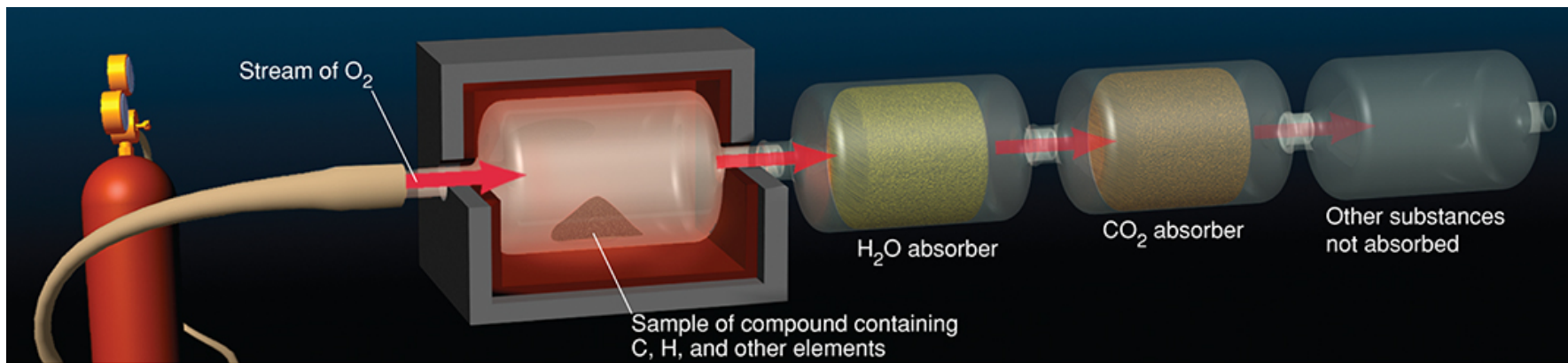
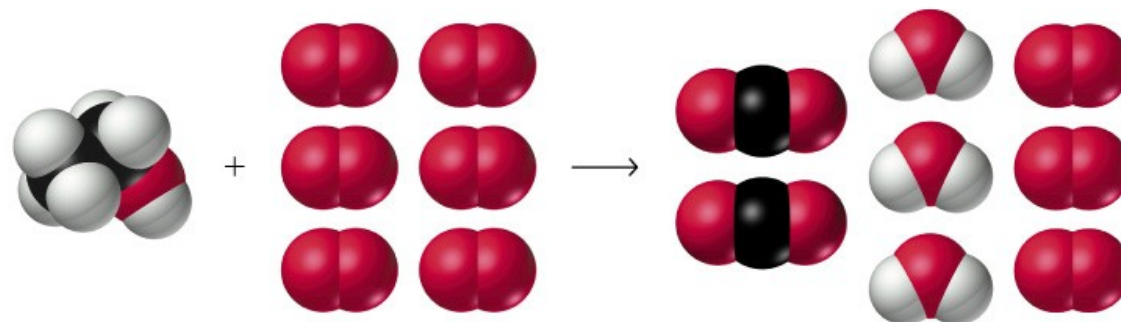
- exact number of atoms of each element

## Structural formula:

- shows relative placement and connectivity of atoms



# Determination of Empirical Formula: Combustion Analysis



# Combustion Data

- the data obtained from combustion analysis of a sample:
  - elements present
  - relative proportions



*empirical  
formula*



# Example 1: Using Combustion Data

Complete combustion of a 1.505 g sample of an unknown compound consisting of C, H and S yields 3.149 g  $\text{CO}_2$ , 0.645 g  $\text{H}_2\text{O}$  and 1.146 g of  $\text{SO}_2$ . What is the empirical formula for the unknown?



# Example 2: Using Combustion Data

When an unknown compound is decomposed into its constituent elements, it is found to contain 71.65% Cl, 24.27% C, and 4.07% H by mass. What is the empirical formula for the unknown?



## Example 2 continued...

We experimentally determine the molecular mass to be 98.96 amu. Determine the molecular formula of the unknown.

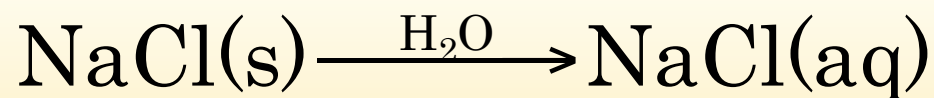


# 3.3 Chemical Equations

- chemical equations depict the kind of reactants and products and their relative amounts in a reaction
- the physical state of the reactants and products may be indicated:



- for the dissolution of compound in water, the following notation may be used:



# Chemical Equations

- Because of the principle of the **conservation of matter**, an equation must be balanced.
- It must have the same number of atoms of the same kind on both sides.



Lavoisier, 1788

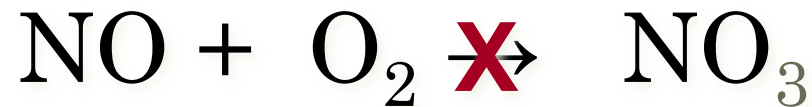


# Balancing Equations: Two Rules

1. Never introduce extraneous atoms to balance.



2. Never change a formula for the purpose of balancing an equation.



## 3.4 Stoichiometry

- the study of the quantitative aspects of chemical reactions
- if the quantity of a reactant is known, it is possible to calculate the quantity of product that will be formed (or vice-versa)
- to simplify the process, we use the mole method: we use units of moles instead of grams or millilitres

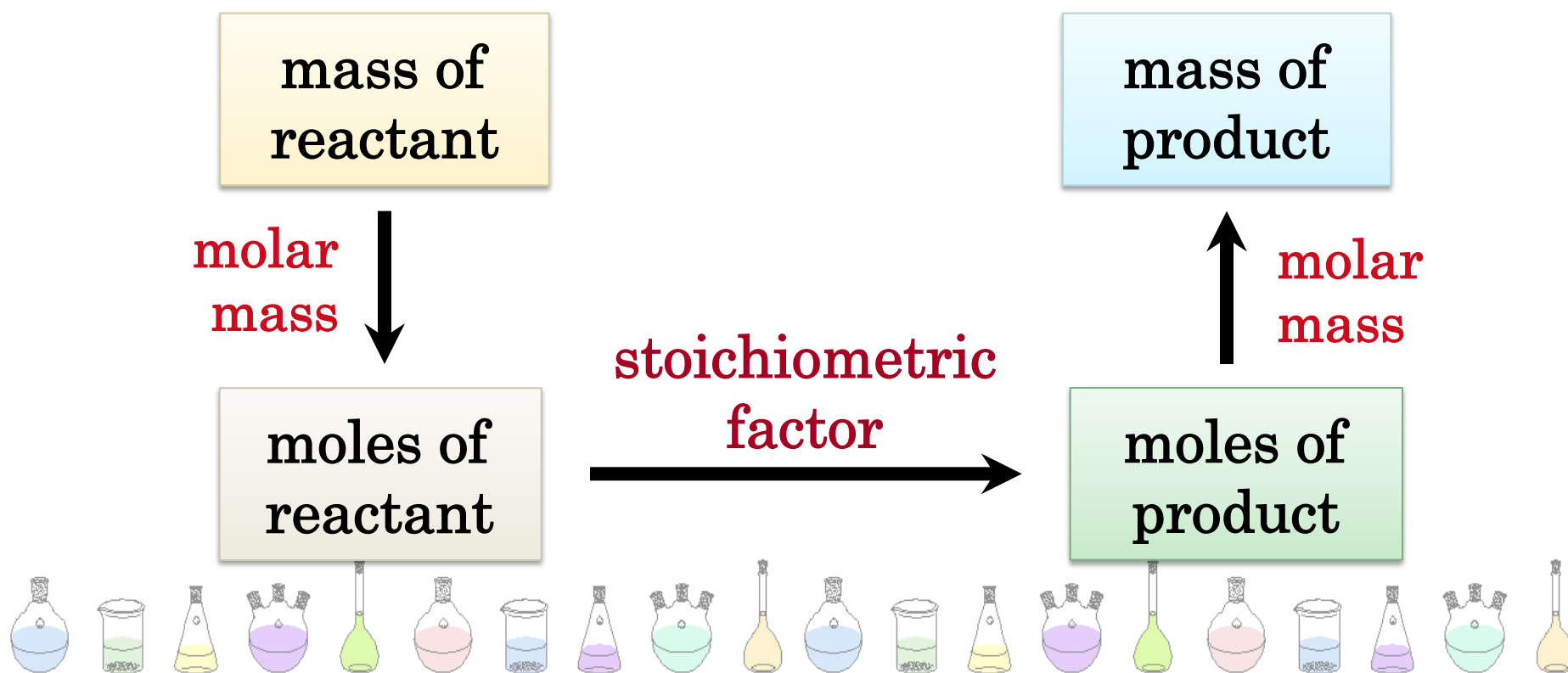


# Using Stoichiometry

Step 1: Identify all reactants and products

Step 2: Balance the chemical equation

Step 3: Follow the mole method



# Your Turn...

What amount (mol) of oxygen atoms is there in 6 moles of  $\text{Al}_2(\text{SO}_4)_3$ ?



# Example: Using the mole method

If 454 g of  $\text{NH}_4\text{NO}_3$  decomposes, how much  $\text{N}_2\text{O}$  and  $\text{H}_2\text{O}$  (in grams) are formed?



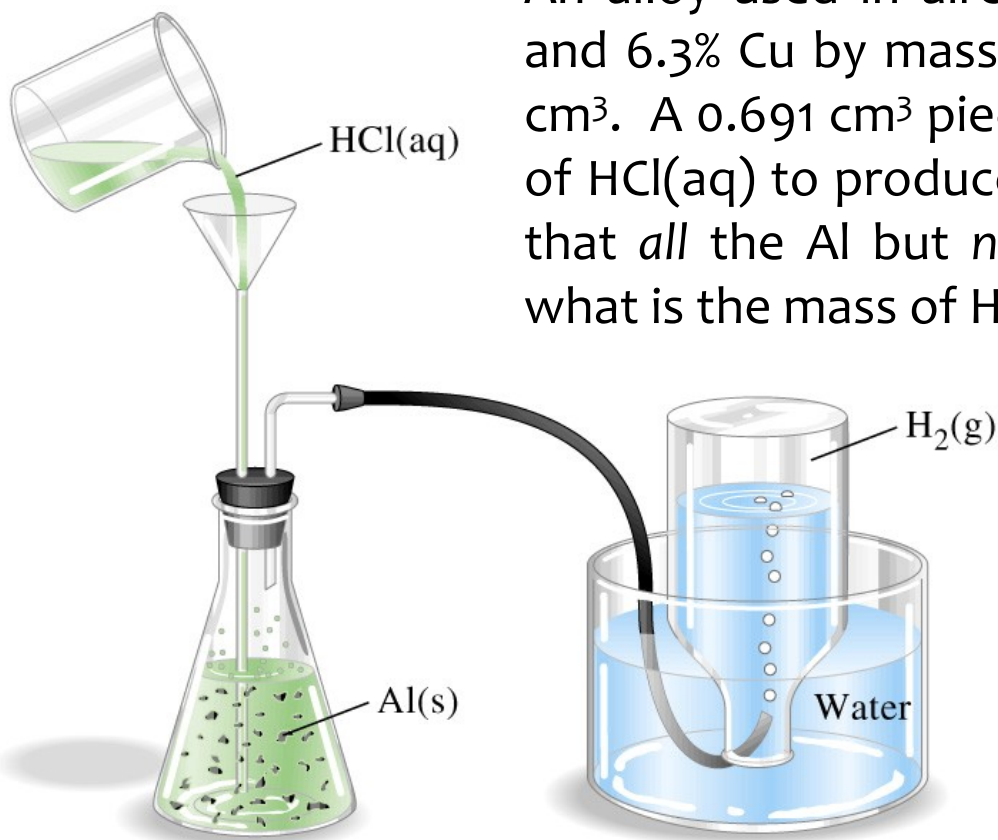
# Stoichiometry Problems

- stoichiometry and the mole method is an extremely powerful tool
- it can be used with other conversion factors, such as:
  - density
  - concentration
  - percent composition



# Example: Using stoichiometry with additional conversion factors

An alloy used in aircraft structures consists of 93.7% Al and 6.3% Cu by mass. The alloy has a density of  $2.85 \text{ g/cm}^3$ . A  $0.691 \text{ cm}^3$  piece of the alloy reacts with an excess of  $\text{HCl(aq)}$  to produce  $\text{H}_2(\text{g})$  and  $\text{AlCl}_3(\text{aq})$ . If we assume that *all* the Al but *none* of the Cu reacts with  $\text{HCl(aq)}$ , what is the mass of  $\text{H}_2$  obtained?



# Reactions Involving a LIMITING REACTANT

- Definition: In a given reaction, there is **not enough of one reagent** to use up the other reagent completely
- The reagent in short supply **LIMITS** the quantity of **product** that can be formed
- The stoichiometric coefficients are used to determine the **limiting reagent**

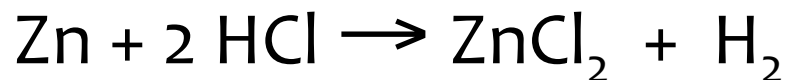


# Limiting Reactants



React solid Zn with 0.100 mol HCl (aq)

The reaction:



	Rxn 1	Rxn 2	Rxn 3
mass Zn (g)	7.00	3.27	1.31
mol Zn	0.107	0.050	0.020
mol HCl	0.100	0.100	0.100
mol HCl/mol Zn	0.93	2.00	5.00



# Example: Limiting Reagent Problem

Given the following reaction:



If 5.40 g of Al are reacted with 8.10 g of  $\text{Cl}_2$ , how many grams of  $\text{Al}_2\text{Cl}_6$  can form?



# Same example, continued...

How much of the reactant in excess will remain when reaction is complete?



# Reaction Yields

- the theoretical yield is the amount of product expected if the reactants react to completion
- the actual yield is always **smaller** than this value!
  - the inverse reaction occurs
  - other side-products are products
  - difficult to collect all of the product

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



# Example: Reaction yields

For the preceding example, if only 7.8 g of  $\text{Al}_2\text{Cl}_6$  were formed, what is the yield of the reaction?



# Your Turn...

Two sets of successive reactions,  $A \rightarrow B$  and  $B \rightarrow C$ , have respective yields of 48% and 73%.

What is the overall percent yield for conversion of A to C?

- A) 66%
- B) 48%
- C) 73%
- D) 35%
- E) I'm not sure.



# 3.5 Chemical Reactions in Solution

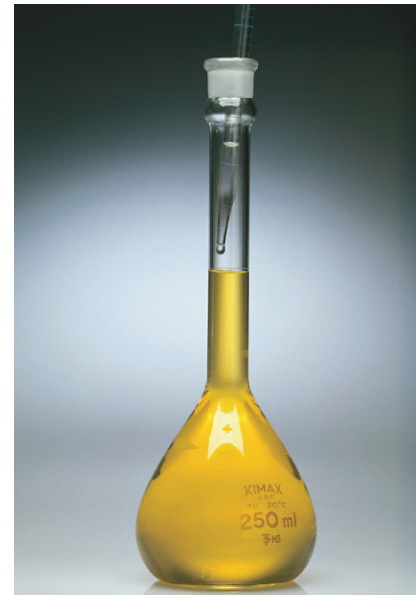
- **Solution:** solute dissolved in solvent to form a homogeneous mixture
- **Solute:** present in smallest amount
- **Solvent:** present in largest amount
  - when the solvent is water = aqueous solutions
  
- **Concentration:** the measure of the amount of solute in a solution
- **Molarity:** Moles of solute per litre of solution



# Concentrations of Solutions

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{volume of solution in liters}}$$

If we know molarity and liters of solution, we can calculate moles (and mass) of solute.



# Example: Solution Concentrations

We want to prepare exactly 0.2500 L of an 0.250 M  $\text{K}_2\text{CrO}_4$  solution in water. What mass of  $\text{K}_2\text{CrO}_4$  should we use?



# 12.4 Other forms of concentration

Molarity	M (or C)	$\frac{n_{\text{solute}}}{V_{\text{solution}}}$
Mass Percent	m/m%	$\frac{m_{\text{solute}}}{m_{\text{solution}}} = \frac{x \text{ g solute}}{100 \text{ g solution}}$
Volume percent	v/v%	$\frac{V_{\text{solute}}}{V_{\text{solution}}} = \frac{x \text{ mL solute}}{100 \text{ mL solution}}$
Mass/volume percent	m/v%	$\frac{m_{\text{solute}}}{V_{\text{solution}}} = \frac{x \text{ g solute}}{100 \text{ mL solution}}$



# Your Turn...

An aqueous solution of hydrogen peroxide,  $\text{H}_2\text{O}_2$ , has a mass/volume percent of 3%. What is the concentration in mol/L?

- A. 0.03 mol/L
- B. 3 mol/L
- C. 0.88 mol/L
- D. 0.088 mol/L
- E. I'm not sure

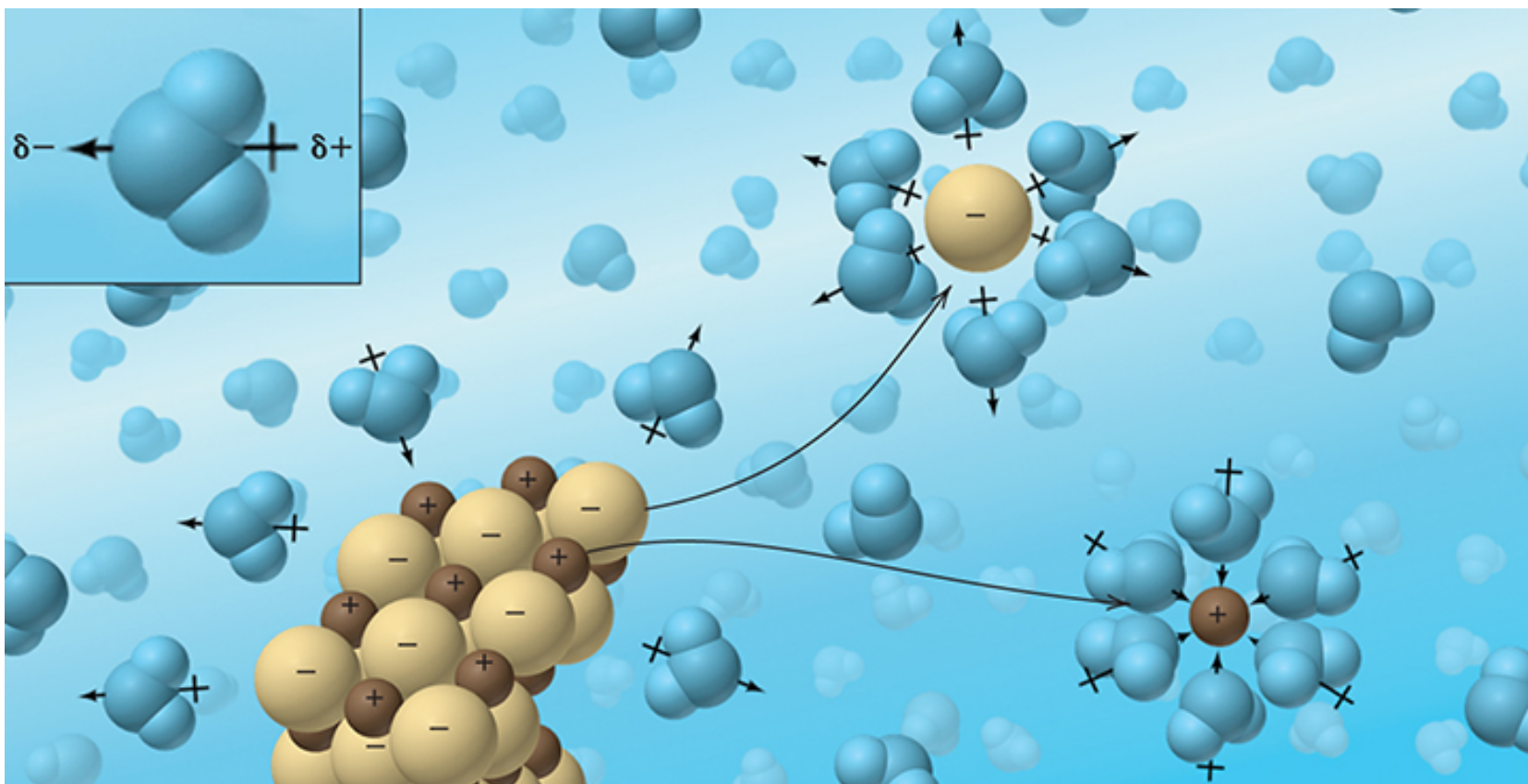


# 12.4 Other forms of concentration

Parts per million	ppm	$\frac{1 \text{ part solute}}{10^6 \text{ parts solution}} = \frac{x \text{ g solute}}{10^6 \text{ g solution}}$ $\approx \frac{x \text{ mg solute}}{\text{L solution}}$
Molality	$m$	$\frac{n_{\text{solute}}}{m_{\text{solvent}}} = \frac{x \text{ mol solute}}{1 \text{ kg solvent}}$
Mole Fraction	$\chi$	$\frac{n_A}{n_{\text{total}}} = \frac{x \text{ mol of "A"}}{1 \text{ mol solution}}$



# 3.6 Water as a solvent



# Electrolytes

The diagram is enclosed in a blue border and consists of three vertical panels, each with a photograph of a light bulb in a beaker at the top and a corresponding circular inset below. Panel A shows a light bulb in a beaker of distilled water, which is unlit. The inset shows a circular view of the water with blue water molecules (two small spheres and one larger sphere) and no ions. Panel B shows a light bulb in a beaker of a white solid, which is unlit. The inset shows a circular view of a solid lattice of green and brown spheres, with '+' and '-' signs indicating fixed ions. Panel C shows a light bulb in a beaker of a clear liquid, which is lit. The inset shows a circular view of the liquid with blue water molecules and green and brown ions. Arrows indicate the movement of positive ions (+) towards the left electrode (labeled 'To (+) electrode') and negative ions (-) towards the right electrode (labeled 'To (-) electrode').

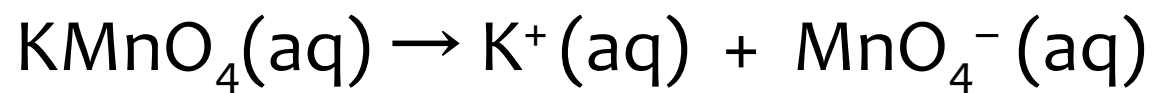
**A** Distilled water does not conduct a current.

**B** Positive and negative ions fixed in a solid do not conduct a current.

**C** In solution, positive and negative ions move and conduct a current.

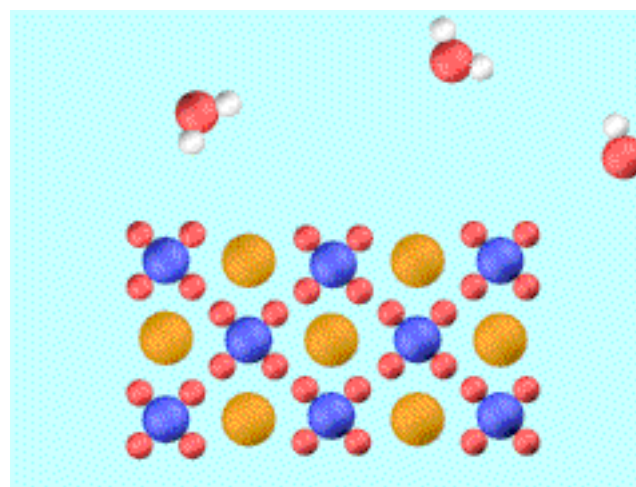


# The Nature of Solutions (The example of $\text{KMnO}_4$ )



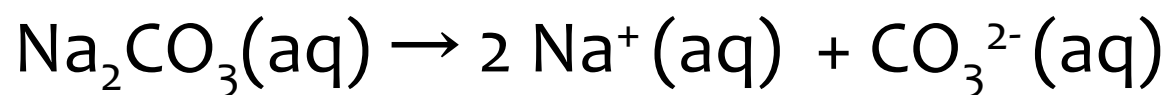
If you make a solution that is 0.30 M in  $\text{KMnO}_4$ , this means that

$$[\text{K}^+] = [\text{MnO}_4^-] =$$



# The Nature of a $\text{Na}_2\text{CO}_3$ Solution

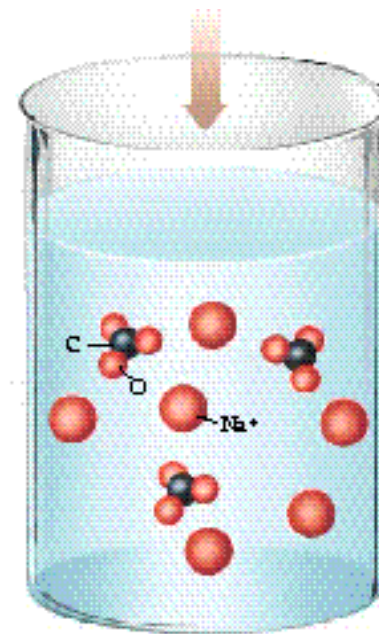
This water-soluble compound is a strong electrolyte:



If  $[\text{Na}_2\text{CO}_3] = 0.100 \text{ M}$ , then

$[\text{Na}^+] =$

$[\text{CO}_3^{2-}] =$



# Your Turn...

In an environmental test for lead (207.2 g/mol) in soil, the lead is oxidized using permanganate ions in acidic solution, according to the following (balanced) reaction:



A 25.0 g sample of soil was found to react with 14.92 mL of a 0.150 M  $\text{KMnO}_4$  solution. What was the percent composition by mass of lead in the soil sample?

- A. 4.6%
- B. 1.2%
- C. 7.4%
- D. 0.46%
- E. I'm not sure

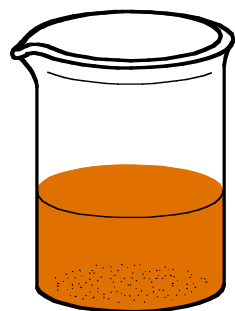


# Concentrations of Solutions

## Dilution

- addition of extra solvent does not change the amount of dissolved solute, therefore:

$$C_{\text{concentrated}} V_{\text{concentrated}} = \text{moles} = \underline{\hspace{10em}}$$



**Add solvent**



**Concentrated**

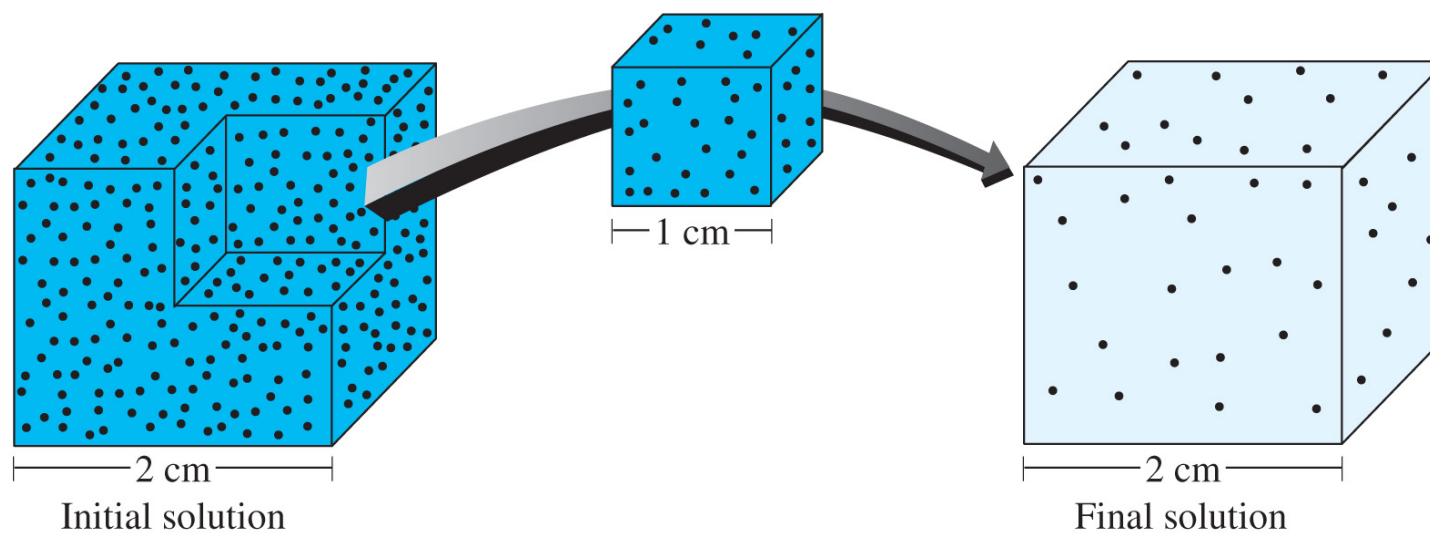
**Dilute**



# Preparing Solutions by Dilution

A handy shortcut:

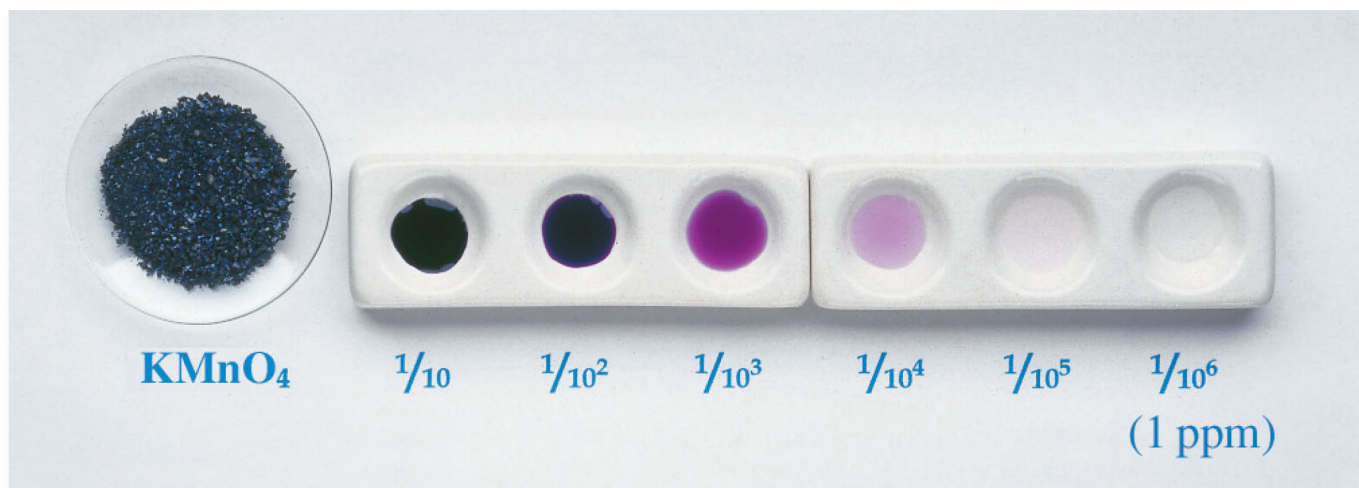
$$C_{\text{initial}} \cdot V_{\text{initial}} = C_{\text{final}} \cdot V_{\text{final}}$$



# Your Turn...

You have a standard solution of 1.00 mol/L  $\text{KMnO}_4$ . To prepare a solution of 100 mL of 0.100 mol/L  $\text{KMnO}_4$ , how many mL of the standard solution will you need?

- A. 0.10 mL
- B. 1 mL
- C. 10 mL
- D. 100 mL



# Homeopathy = dilutions!

- Based on the principle of “similarities”
- Belief: the greater the dilution, the more powerful the remedy (!?)

1 CH	$0,01 = 10^{-2}$	10 mL in 1 L
2 CH	$0,0001 = 10^{-4}$	1 mL in 10 L
7 CH	$10^{-14}$	1 L in a large lake
15 CH	$10^{-30}$	one molecule in 30 tonnes of water
30 CH	$10^{-60}$	1 L in the entire galaxy
40 CH	$10^{-80}$	one molecule in the whole universe
200 CH	$10^{-400}$	homeopathic flu remedy: <i>Oscillococcinum</i>



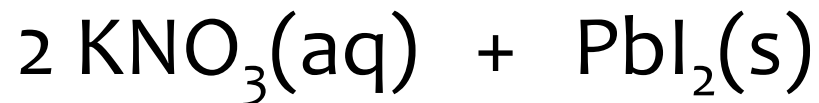
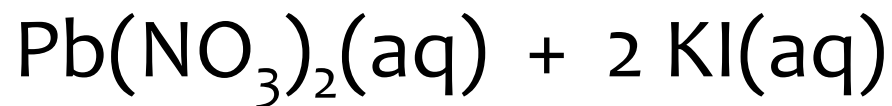
# Your Turn... A challenge!

“Weak” homeopathic remedies are made from 13CH (or 26X) dilutions. This corresponds to a ratio of **1 g of substance to  $10^{26}$  g of solution**. Let’s say a remedy is prepared from a substance with MM = **100 g/mol** – how many litres of this formulation must I drink in order to consume **one molecule** of the active ingredient? Assume the solution has the same density as water, **1.0 g/mL**.



# 3.7 Aqueous ionic reactions

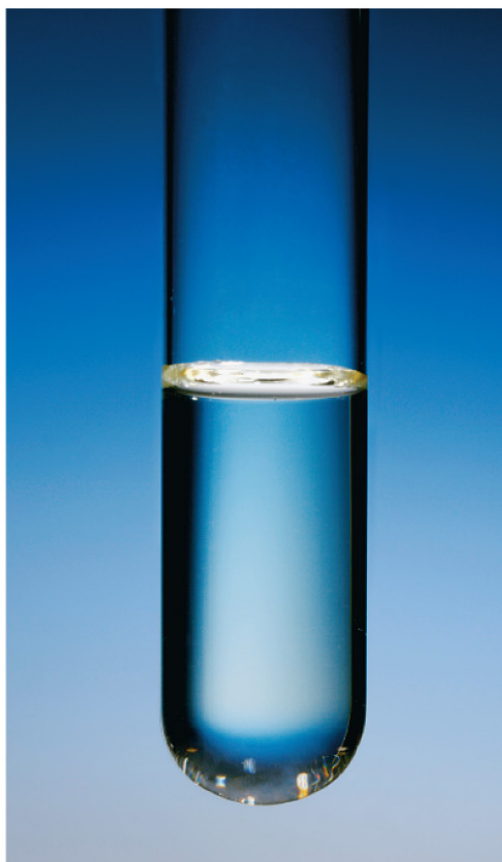
- The “driving force” is the formation of a solid or gas:



# An example: $\text{AgNO}_3 + \text{NaI}$



$\text{AgNO}_3(\text{aq})$



$\text{NaI}(\text{aq})$

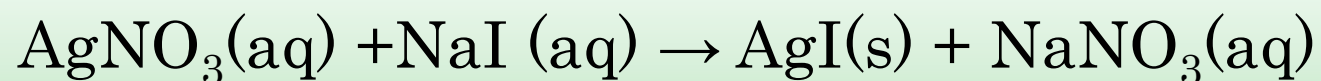


$\text{AgI}(\text{s})$



# Writing Aqueous Ionic Reactions

## A. Molecular equation

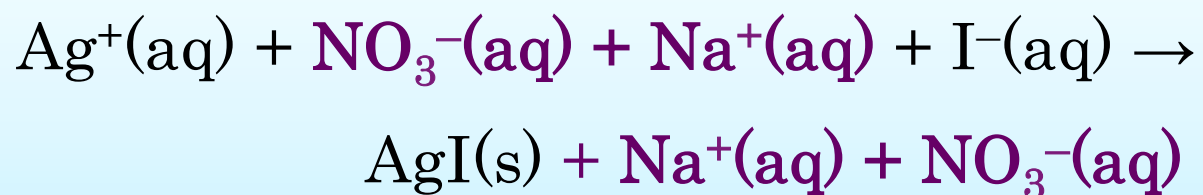


- species are written as intact compounds
- in reality: they are dissociated ions!



# Writing Aqueous Ionic Reactions

## B. Ionic equation

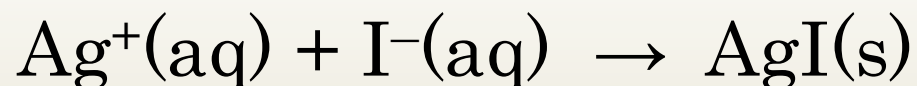


- more realistic: represents all species in their ionic forms
- in this example,  $\text{Na}^+$  and  $\text{NO}_3^-$  are known as \_\_\_\_\_



# Writing Aqueous Ionic Reactions

## C. Net ionic equation

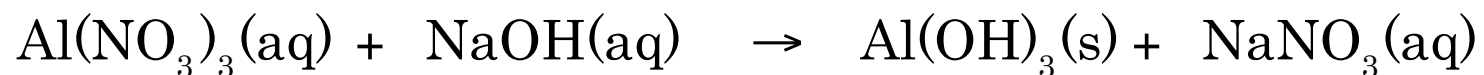


- this equation drops all of the spectator ions, and so includes only those ions participating in the reaction



# Example: Aqueous Ionic Reactions

Write and balance the net ionic equation.



# 19.1 Redox Reactions

- an oxidation-reduction, or REDOX, reaction involves the transfer of electrons
- every redox reaction contains a reducing agent and an oxidizing agent

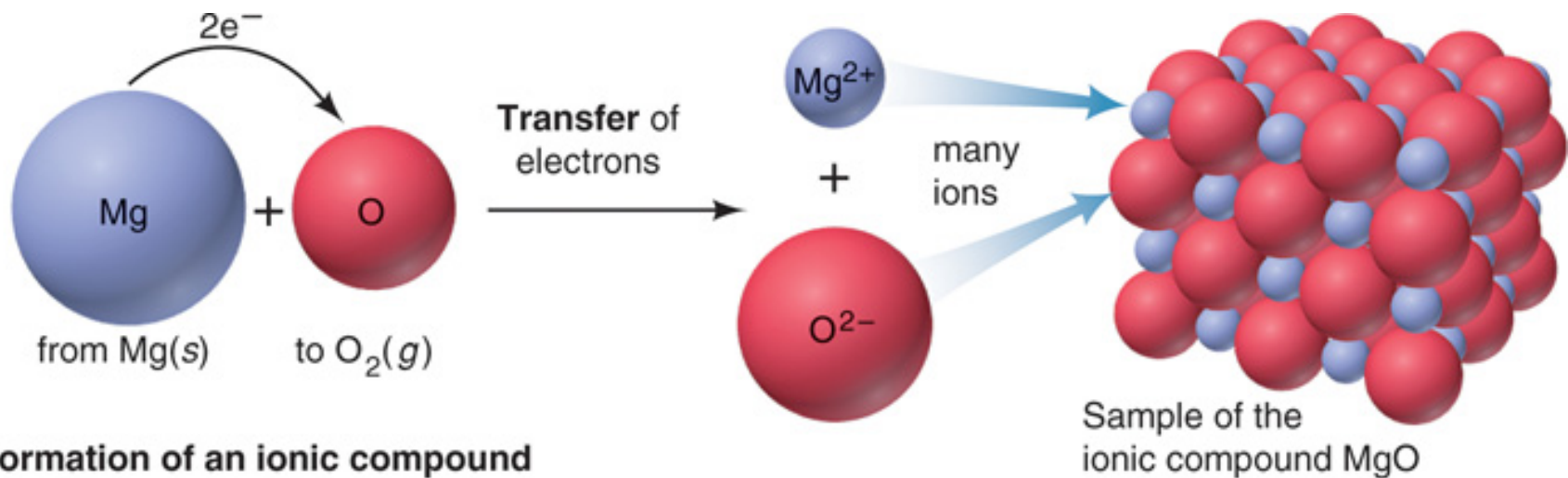
– the reducing agent \_\_\_\_\_ electrons

– the oxidizing agent \_\_\_\_\_ electrons



# Redox Reactions

If something has been oxidized then something else has also been reduced.



# Assigning Oxidation States, p. 787

**TABLE 19.1** Rules for Assigning an Oxidation Number

## General Rules

1. For an atom in its elemental form (such as Na, O<sub>2</sub>, and Cl<sub>2</sub>): Oxidation number = 0
2. For a monatomic ion: Oxidation number = ion charge (with the sign *before* the numeral)
3. The sum of the oxidation number values for the atoms in a molecule or formula unit of a compound equals zero. The sum of the oxidation number values for the atoms in a polyatomic ion equals the charge of the ion.

## Rules for Specific Atoms or Periodic Table Groups

1. For Group 1: Oxidation number = +1 in all compounds
2. For Group 2: Oxidation number = +2 in all compounds
3. For hydrogen: Oxidation number = +1 in combination with nonmetals  
Oxidation number = -1 in combination with metals and boron
4. For fluorine: Oxidation number = -1 in all compounds
5. For oxygen: Oxidation number = -2 in most cases, unless coupled to a more electronegative centre (such as F) or a Group 1 or Group 2 metal (in which case, it might be +2 or -1)
6. For Group 17: Oxidation number = -1 in combination with metals, nonmetals (except O), and other halogens lower in the group



# Some Tips...

- if two rules contradict each other, follow the rule that appears higher in the list.
- the sum of the oxidation states of all the atoms must be equal to the charge of the ion or molecule
- for a multi-atom species, label the easy oxidation states, and solve for the unknown atoms



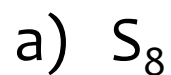
# Example: Deriving oxidation states

Label the oxidation states of each element in  $\text{MnO}_4^-$ .

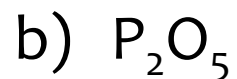


# More examples...

What is the oxidation state of the elements in each of the following?



S =



P =

O =



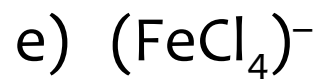
Cr =

O =



Ca =

H =



Fe =

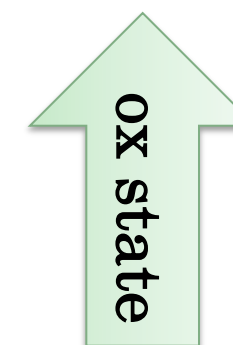
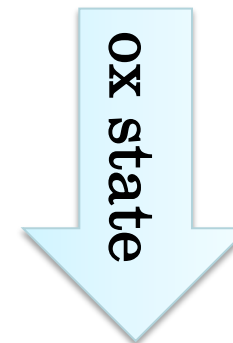
Cl =



# Oxidation-Reduction Reactions

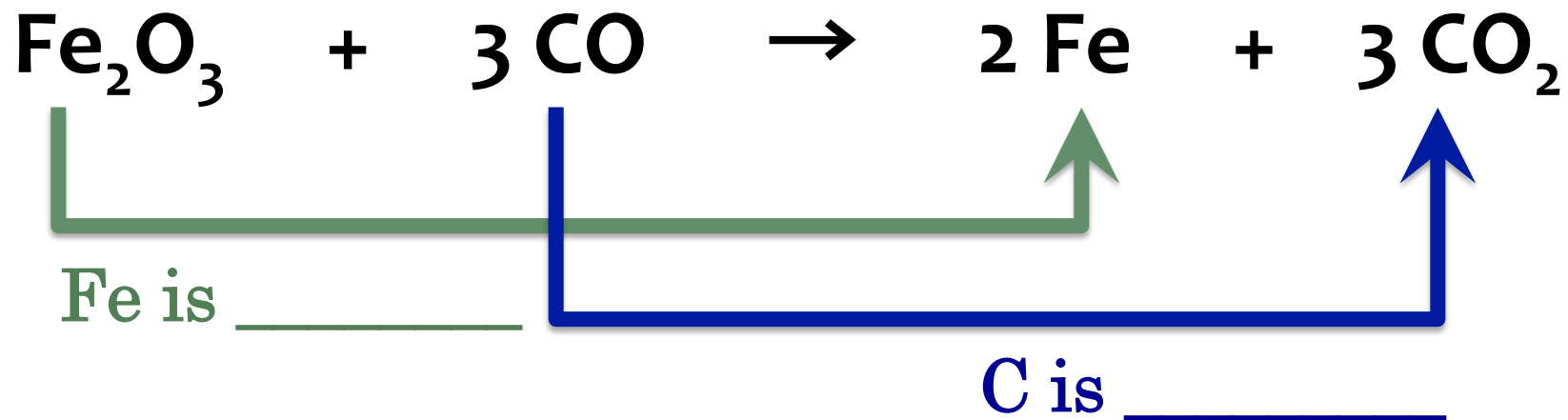
Some terminology:

- An **oxidizing agent** (oxidant):
  - oxidation state decreases in a redox reaction (reduced)
- A **reducing agent** (reductant):
  - oxidation state increases in a redox reaction (oxidized)



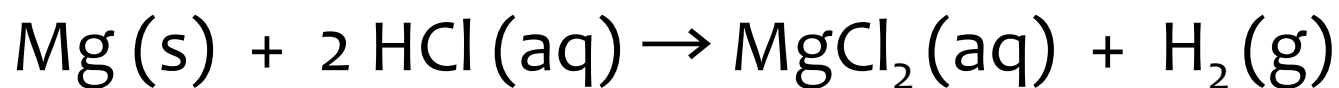
# Redox Reactions

We can determine which species is oxidized and which species is reduced by looking at the oxidation states (Chapter 3)



# Your Turn...

In the following reaction, which is the reducing agent?



- A. Mg
- B. HCl
- C.  $\text{MgCl}_2$
- D.  $\text{H}_2$
- E. I'm not sure



# Redox Example: Zinc in Copper Sulfate

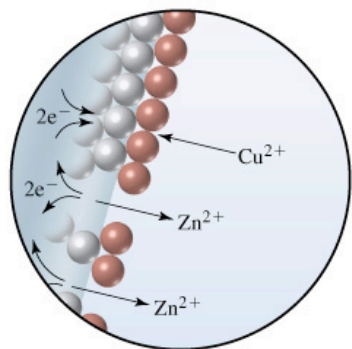


(a)



(b)

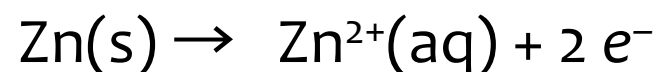
Oxidation-Reduction  
Reactions - Part 2



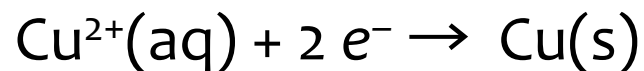
# Half Reactions

A redox reaction is represented by two half-reactions.

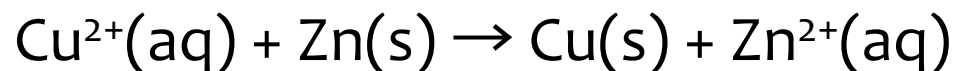
**Oxidation:**



**Reduction:**



**Overall:**



Balancing redox reactions depends on whether it is an *acidic* or *basic* medium...



# Balancing Redox Reactions

## The Half-Equation Method:

- the overall reaction is separated into two half-reactions: an **oxidation** and a **reduction**
- each half-reaction is balanced
- the balanced half-reactions are then added to give the overall reaction



# The Half-Reaction Method (in acidic medium)

**STEP (1):** Write out the net ionic form of the reaction.

**STEP (2):** Separate into two half-reactions.

**STEP (3):** Balance all elements except O and H in the half-reactions.

**STEP (4):** Add enough molecules of  $\text{H}_2\text{O}$  to balance the O atoms.

**STEP (5):** Add enough  $\text{H}^+$  ions to balance the H atoms.

**STEP (6):** Balance the net charge on each side by adding electrons.

**STEP (7):** If necessary, multiply the half-reactions by the LCD.

**STEP (8):** Add the two half-reactions together and simplify.

**STEP (9):** Verify your final reaction.



# Example: Balancing in acidic solution

Balance the following redox reaction.



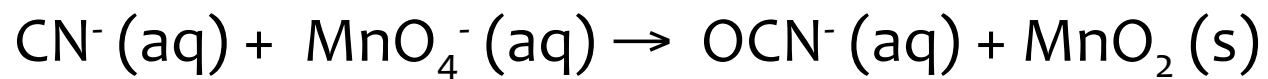
# In basic solution...

- when the reaction is in basic solution, **STEP (5)** is modified as follows:
  - for every  $\text{H}^+$ , add an  $\text{OH}^-$  on each side of the half-reaction
  - where  $\text{H}^+$  and  $\text{OH}^-$  appear on the same side, combine these two ions to make  $\text{H}_2\text{O}$



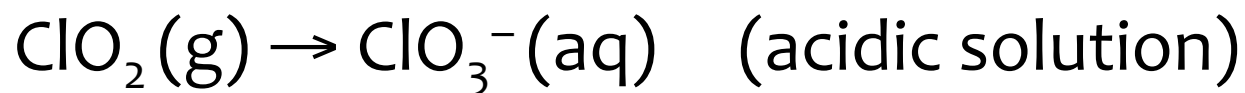
# Example: Balancing in basic solution

Balance the following redox reaction in basic medium.



# Your Turn...

Is the following half-reaction an oxidation or reduction?



- A. oxidation
- B. reduction
- C. I'm not sure



# Chapter 3: Key Concepts

- the mole and Avogadro's number
- percent composition
- empirical vs. molecular formulae
- balancing chemical equations
- stoichiometry and the mole method
- limiting reactant problems
- percent yields
- concentrations of solutions
- aqueous ionic reactions



# Sections 12.4 and 19.1 Key Concepts

- various concentration terms
- oxidation states
- reduction vs. oxidation half-reactions
- oxidizing vs. reducing agents
- balancing redox reactions



# Chapter 3: Suggested Problems

3.2, 3.10, 3.16, 3.23, 3.27, 3.38,  
3.40, 3.44, 3.47, 3.54, 3.60, 3.66,  
3.70, 3.76, 3.78, 3.84, 3.88, 3.99,  
3.101, 3.105, 3.112, 3.123, 3.129,  
3.133, 3.135, 3.141, 3.142, 3.148,  
3.150, 3.155, 3.165, 3.172, 3.176, 3.182



# Section 12.4: Suggested Problems

12.49, 12.53, 12.57,  
12.63, 12.65, 12.69,  
12.119, 12.133, 12.137,  
12.141, 12.152



# Section 19.1: Suggested Problems

19.3, 19.4, 19.10, 19.12,  
19.14, 19.16, 19.21

