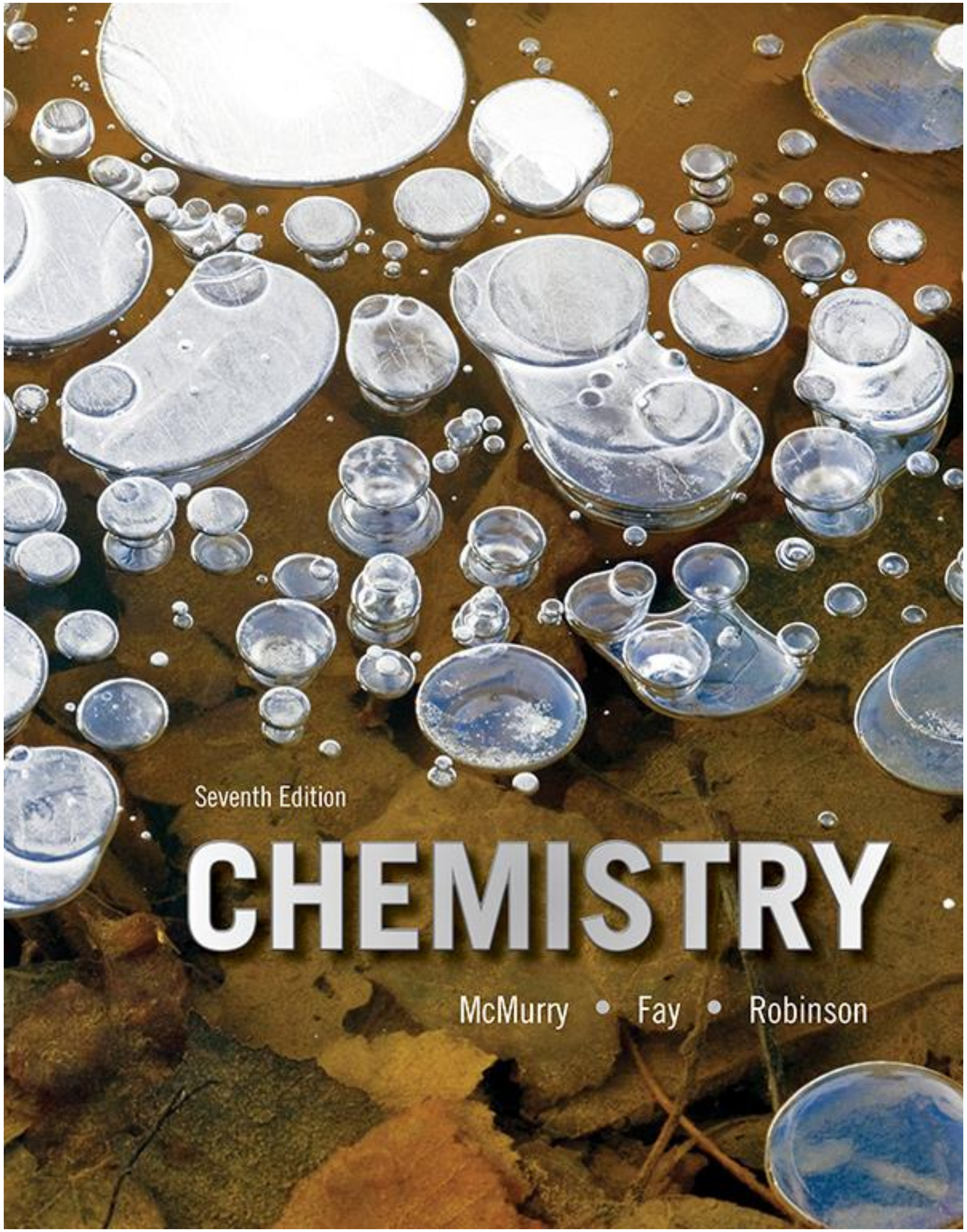


Lecture Presentation

Chapter 4

Reactions in Aqueous Solution

4.1, 4.2, 4.3, 4.4, 4.6, 4.7, 4.8, 4.12, 4.15, 4.18, 4.20, 4.22, 4.23, 4.24, 4.26, 4.28, 4.31, 4.42, 4.48, 4.52, 4.54, 4.64, 4.68, 4.70, 4.72, 4.90, 4.98, 4.104, 4.110, 4.138

John E. McMurry
Robert C. Fay

Solution Concentration: Molarity

Molarity: The number of moles of a substance dissolved in each liter of solution

Solution: A homogeneous mixture

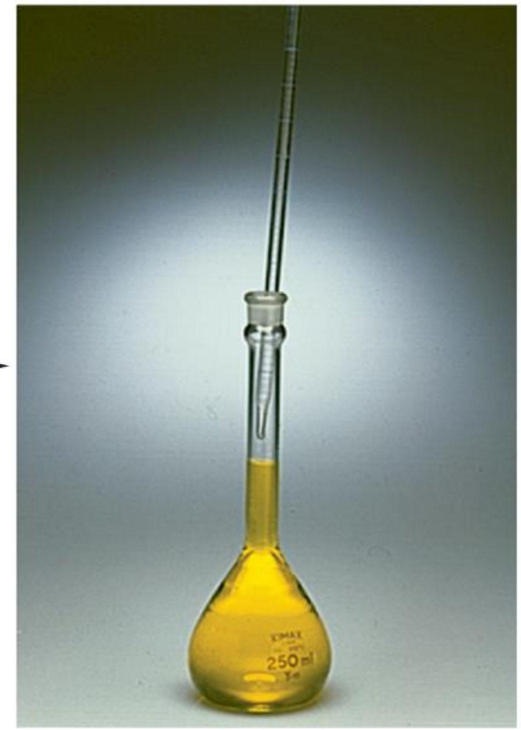
Solute: The dissolved substance in a solution (substance you have less quantity of in solution)

Solvent: The major component in a solution (substance you have more of in solution)

Solution Concentration: Molarity

A measured number of moles of **solute** is placed in a volumetric flask.

Enough solvent is added to dissolve the solute by swirling.



Further solvent is added to reach the calibration mark on the neck of the flask, and the solution is mixed until uniform.

Definition of Molarity

Molarity (*M*)

- Commonly used expression for concentration
- Defined as moles of solute per volume of solution in liters

$$M = \text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$\text{also } M = \frac{\text{moles of solute}}{\text{1000 ml of solution}}$$

← use if given volume of solution in mL

Definition of Molarity

$$M = \text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

Example: What is the molarity if you put 1.00 mol of sodium chloride in enough water to make 1.00 L of solution ? End class D section 9/30

$$\frac{1.00 \text{ mol}}{1.00 \text{ L}} = 1.00 \frac{\text{mol}}{\text{L}} \text{ or } 1.00 \text{ M}$$

End class D section 9/30

Solution Concentration: Molarity

- Calculate the molarity of a solution prepared by dissolving 1.56 g of gaseous HCl (36.46 g/mol) in enough water to make 26.8 mL of solution

- Doc camera

Solution Concentration: Molarity

- Where are we going?
  - To find the molarity of HCl solution
  - What do we know?
    - 1.56 g HCl
    - 26.8 mL solution
  - What information do we need to find molarity?
    - Moles solute

$$\text{Molarity} = \frac{\text{mol solute}}{\text{L solution}}$$

Solution Concentration: Molarity

- How do we get there?
  - What are the moles of HCl (36.46 g/mol)?

$$1.56\ \cancel{\text{g HCl}} \times \frac{1\ \text{mol HCl}}{36.46\ \cancel{\text{g HCl}}} = 4.28\times10^{-2}\ \text{mol HCl}$$

  - What is the volume of solution (in liters)?

$$26.8\ \cancel{\text{mL}} \times \frac{1\ \text{L}}{1000\ \cancel{\text{mL}}} = 2.68\times10^{-2}\ \text{L}$$

Solution Concentration: Molarity

  - What is the molarity of the solution?

$$\text{Molarity} = \frac{4.28 \times 10^{-2}\ \text{mol HCl}}{2.68 \times 10^{-2}\ \text{L solution}} = 1.60\ M\ \text{HCl}$$

- Reality check
  - The units are correct for molarity

Solution Concentration: Molarity

Determining Moles of Solute in a Sample

- Use the definition of molarity as a conversion factor (to get moles of solute)

$$\text{Liters of solution} \times \text{molarity} = \cancel{\text{liters of solution}} \times \frac{\text{moles of solute}}{\cancel{\text{liters of solution}}}$$

$$\boxed{\text{Moles of solute} = \text{Liters of solution} \times \text{molarity}}$$

Solution Concentration: Molarity

- Calculate the number of moles of Cl^{-} ions in 1.75 L of 0.152 *M Na*Cl   (M = moles / liter solution)
- On doc camera

HW: Solution Concentration: Molarity

How many grams of solute would you use to prepare 1.50 L of 0.250 M glucose, $C_6H_{12}O_6$?

Molar mass $C_6H_{12}O_6$ = 180.0 g/mol

Solution Concentration: Molarity

How many grams of solute would you use to prepare 1.50 L of 0.250 M glucose, $C_6H_{12}O_6$?

Molar mass $C_6H_{12}O_6$ = 180.0 g/mol

$$1.50\ \text{L} \times \frac{0.250\ \text{mol}}{1\ \text{L}} \times \frac{180.0\ \text{g}}{1\ \text{mol}} = \boxed{67.5\ \text{g}}$$

Dilution

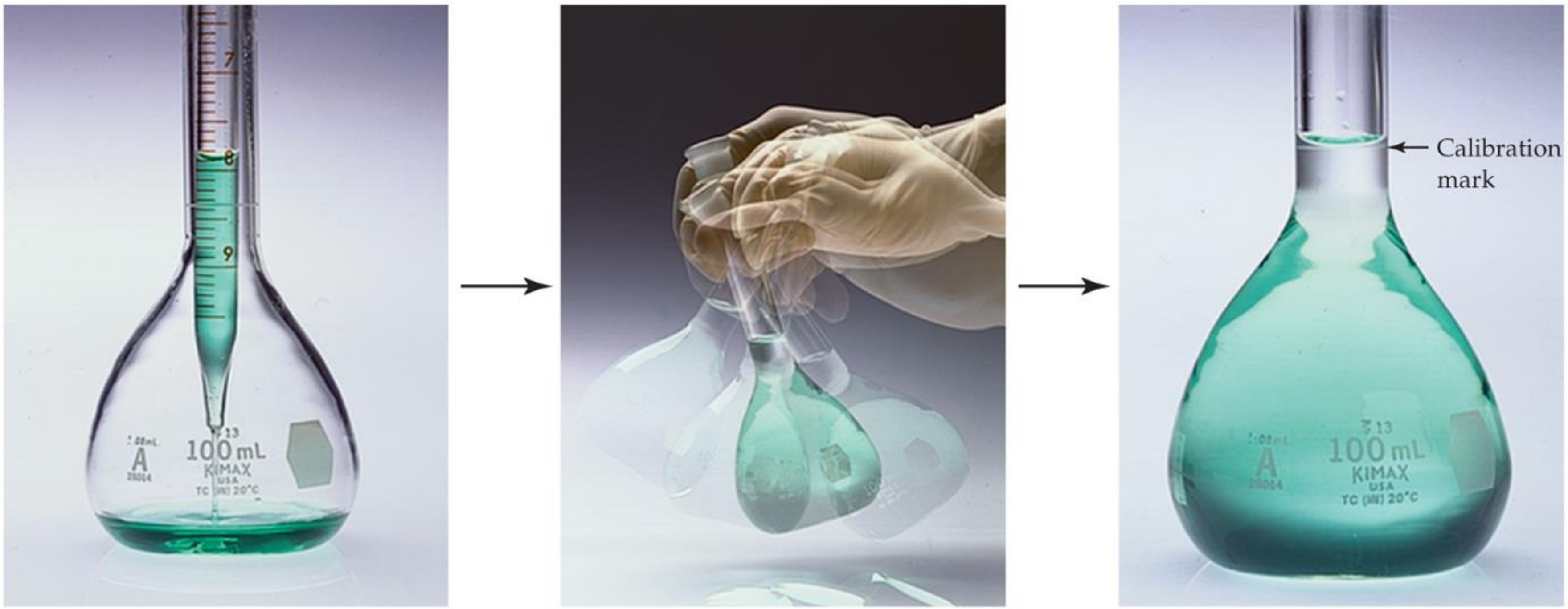
- Process of adding water to a concentrated (stock) solution

moles of solute before dilution = moles of solute after dilution

$M_i V_i$ = moles initial , $M_f V_f$ = moles final

- dilution equation: $M_i V_i = M_f V_f$
- as long as keep volume unit the same between initial and final, can have both mL or both L

Diluting Concentrated Solutions



The **volume** to be diluted is placed in an empty volumetric flask.

Solvent is added to a level just below the calibration mark, and the flask is shaken.

More solvent is added to reach the calibration mark, and the flask is again shaken.

Diluting Concentrated Solutions

How would you prepare 250.0 mL of 0.500 M from 18.0 M aqueous H_2SO_4?

$M_i = 18.0$ M $M_f = 0.500$ M

$V_i = ?$ mL $V_f = 250.0$ mL

$$V_i = \frac{M_f V_f}{M_i} = \frac{0.500\text{ M}}{18.0\text{ M}} \times 250.0\text{ mL} = \boxed{6.94\text{ mL}}$$

Add 6.94 mL 18.0 M sulfuric acid to enough water to make 250.0 mL of 0.500 M solution.

HW: Diluting Concentrated Solutions [do (a) only for HW]

- Describe how you would prepare 2.00 L of each of the following solutions: (on doc camera)
  a. 0.250 *M* NaOH from 1.00 *M* NaOH stock solution

  b. 0.100 *M* K_2CrO_4 from 1.75 *M* K_2CrO_4 stock solution

HW: Diluting Concentrated Solutions [do (a) only for HW]

End 9/30 F section

- Describe how you would prepare 2.00 L of each of the following solutions: (on doc camera)
  a. 0.250 *M* NaOH from 1.00 *M* NaOH stock solution

     Add 500. mL (= 0.500 L) of the 1.00 *M* NaOH stock solution to a 2-L volumetric flask; fill to the mark with water

  b. 0.100 *M* K_2CrO_4 from 1.75 *M* K_2CrO_4 stock solution

     Add 114 mL (= 0.114 L) of the 1.75 *M* K_2CrO_4 stock solution to a 2- L volumetric flask; fill to the mark with water

Electrolytes in Aqueous Solution

Electrolytes: dissociate in water to produce conducting solutions of ions

$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

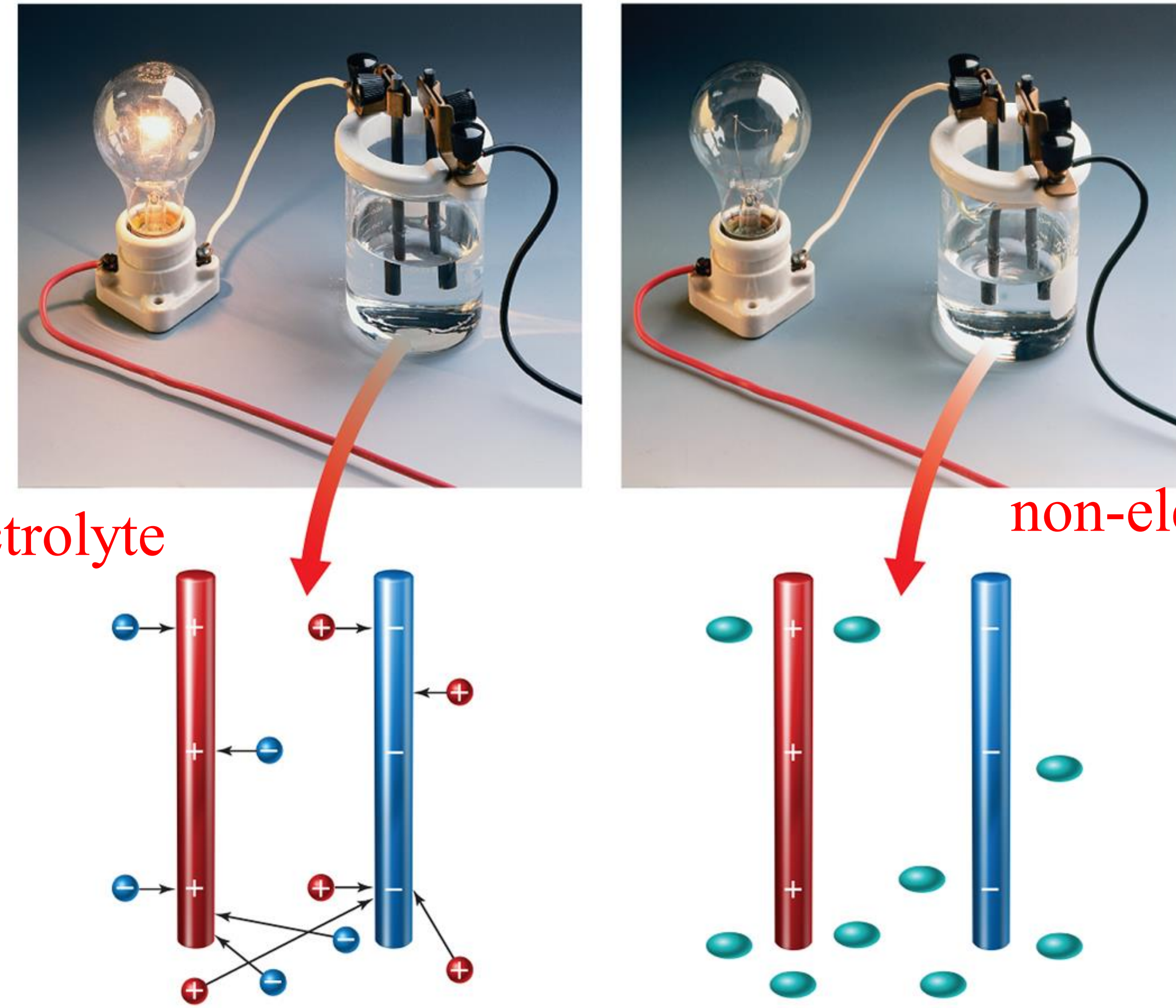
Nonelectrolytes: do not dissociate to produce ions in water (covalent compounds)

$$C_{12}H_{22}O_{11}(s) \xrightarrow{H_2O} C_{12}H_{22}O_{11}(aq)$$

Electrolytes in Aqueous Solution

electrolyte

non-electrolyte



A solution of NaCl conducts electricity because of the movement of charged particles (ions), thereby completing the circuit and allowing the bulb to light.

A solution of sucrose does not conduct electricity or complete the circuit because it contains no mobile charged particles. The bulb therefore remains dark.

Electrolytes in Aqueous Solution

Strong Electrolytes: completely dissociates in water (ionic compounds, strong acids, strong bases)

$$KCl(aq) \longrightarrow K^{+}(aq) + Cl^{-}(aq)$$

Weak Electrolytes: incompletely dissociates in water (weak acids, weak bases)

$$CH_3CO_2H(aq) \rightleftharpoons H^{+}(aq) + CH_3CO_2^{-}(aq)$$

Electrolytes in Aqueous Solution

TABLE 4.1 Electrolyte Classification of Some Common Substances

Strong Electrolytes	Weak Electrolytes	Nonelectrolytes
HCl, HBr, HI	CH_3CO_2H	H_2O
$HClO_4$	HF	CH_3OH (methyl alcohol)
HNO_3	HCN	C_2H_5OH (ethyl alcohol)
H_2SO_4		$C_{12}H_{22}O_{11}$ (sucrose)
KBr		Most compounds of carbon (organic compounds)
NaCl		
NaOH, KOH		
Other soluble ionic compounds		

Strong Acids: Hydrochloric acid, hydrobromic acid, hydroiodic acid, perchloric acid, nitric acid, sulfuric acid

Electrolytes in Aqueous Solution

TABLE 4.1 Electrolyte Classification of Some Common Substances

Strong Electrolytes	Weak Electrolytes	Nonelectrolytes
HCl, HBr, HI	CH_3CO_2H	H_2O
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HNO_3	HCN	C_2H_5OH (ethyl alcohol)
H_2SO_4		$C_{12}H_{22}O_{11}$ (sucrose)
KBr		Most compounds of carbon (organic compounds)
NaCl		
NaOH, KOH		
Other soluble ionic compounds		

Ionic compounds

Electrolytes in Aqueous Solution

TABLE 4.1 Electrolyte Classification of Some Common Substances

Strong Electrolytes	Weak Electrolytes	Nonelectrolytes
HCl, HBr, HI	CH_3CO_2H	H_2O
$HClO_4$	HF	CH_3OH (methyl alcohol)
HNO_3	HCN	C_2H_5OH (ethyl alcohol)
H_2SO_4		$C_{12}H_{22}O_{11}$ (sucrose)
KBr		Most compounds of carbon (organic compounds)
NaCl		
NaOH, KOH		
Other soluble ionic compounds		

Weak acids

Electrolytes in Aqueous Solution

TABLE 4.1 Electrolyte Classification of Some Common Substances

Strong Electrolytes	Weak Electrolytes	Nonelectrolytes
HCl, HBr, HI	CH_3CO_2H	H_2O
$HClO_4$	HF	CH_3OH (methyl alcohol)
HNO_3	HCN	C_2H_5OH (ethyl alcohol)
H_2SO_4		$C_{12}H_{22}O_{11}$ (sucrose)
KBr		Most compounds of carbon (organic compounds)
NaCl		
NaOH, KOH		
Other soluble ionic compounds		

Molecular (covalent) compounds

End class 9/30 Monday G section

Types of Chemical Reactions in Aqueous Solution

Precipitation Reactions: soluble reactants give insoluble solid product (precipitates out)

$$Pb(NO_3)_2(aq) + 2KI(aq) \longrightarrow 2KNO_3(aq) + \boxed{PbI_2(s)}$$

Types of Chemical Reactions in Aqueous Solution

Oxidation–Reduction (Redox) Reactions: one or more electrons (negative charge) are transferred between reaction partners

2 (+ 1) electrons

$$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$$

zero +1 -1 +2 -1 zero

- 2 electrons

Types of Chemical Reactions in Aqueous Solution

Acid-Base Neutralization Reactions: acid reacts with base to give water & salt (salt is ionic compound, not always NaCl)

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

acid	base	water	salt

Aqueous Reactions and Net Ionic Equations (different ways to write reactions)

Molecular Equation: write complete formulas as if *molecules*.

$$Pb(NO_3)_2(aq) + 2KI(aq) \longrightarrow 2KNO_3(aq) + PbI_2(s)$$

Strong electrolytes

Precipitate

Aqueous Reactions and Net Ionic Equations (different ways to write reactions)

Ionic Equation: write strong electrolytes as dissociated ions. (solid, liquid, gas, weak electrolyte compounds written as molecular formula-do NOT dissociate)

$$\overbrace{Pb^{2+}(aq) + 2NO_3^-(aq)}^{Pb(NO_3)_2(aq)} + \overbrace{2K^+(aq) + 2I^-(aq)}^{2KI(aq)}$$

$$\longrightarrow \underbrace{2K^+(aq) + 2NO_3^-(aq)}_{2KNO_3(aq)} + PbI_2(s)$$

Aqueous Reactions and Net Ionic Equations (different ways to write reactions)

Spectator Ions: ions that do not change during the reaction. (& just watch the reaction without doing anything)

$$Pb^{2+}(aq) + 2NO_3^{-}(aq) + 2K^{+}(aq) + 2I^{-}(aq)$$

$$\longrightarrow 2K^{+}(aq) + 2NO_3^{-}(aq) + PbI_2(s)$$

Aqueous Reactions and Net Ionic Equations (different ways to write reactions)

Net Ionic Equation: Only the ions undergoing change are shown. (leave out spectators)

$$Pb^{2+}(aq) + 2I^{-}(aq) \longrightarrow PbI_2(s)$$

End 10/2 W D section

Precipitation Reactions and Solubility Guidelines

Typo L^+ is really Li^+

TABLE 4.2 Solubility Guidelines for Ionic Compounds in Water

Soluble Compounds	Common Exceptions
L^+, Na^+, K^+, Rb^+, Cs^+ (group 1A cations)	None
NH_4^+ (ammonium ion)	None
Cl^-, Br^-, I^- (halide)	Halides of Ag^+, Hg_2^{2+}, Pb^{2+}
NO_3^- (nitrate)	None
ClO_4^- (perchlorate)	None
$CH_3CO_2^-$ (acetate)	None
SO_4^{2-} (sulfate)	Sulfates of Sr^{2+}, Ba^{2+}, Hg_2^{2+}, Pb^{2+}
Insoluble Compounds	**Common Exceptions**
CO_3^{2-} (carbonate)	Carbonates of group 1A cations, NH_4^+
S^{2-} (sulfide)	Sulfides of group 1A cations, NH_4^+, Ca^{2+}, Sr^{2+}, and Ba^{2+}
PO_4^{3-} (phosphate)	Phosphates of group 1A cations, NH_4^+
OH^- (hydroxide)	Hydroxides of group 1A cations, NH_4^+, Ca^{2+}, Sr^{2+}, and Ba^{2+}

HW: Are the following compounds soluble or insoluble in water ?

Li_2CO_3

$BaSO_4$

$Mg(NO_3)_2$

Precipitation Reactions and Solubility Guidelines

Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of $AgNO_3$ and Na_2CO_3 are mixed.

Precipitation Reactions and Solubility Guidelines

Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of $AgNO_3$ and Na_2CO_3 are mixed.

Write the chemical formulas of the products (use proper ionic rules). (exchange cation / anion partners)

$$AgNO_3(aq) + Na_2CO_3(aq) \longrightarrow Ag_2CO_3 + NaNO_3$$

$$AB + CD \longrightarrow CB + AD$$

Cation (on left) Anion (on right)

double replacement reaction

Precipitation Reactions and Solubility Guidelines

Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of $AgNO_3$ and Na_2CO_3 are mixed.

2. **Molecular Equation**: Balance the equation and predict the solubility of each possible product.

$$2AgNO_3(aq) + Na_2CO_3(aq) \longrightarrow Ag_2CO_3(s) + 2NaNO_3(aq)$$

insoluble (look up solubility chart)

contains a group 1A cation (soluble from solubility chart)

Precipitation Reactions and Solubility Guidelines

Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of $AgNO_3$ and Na_2CO_3 are mixed.

3. **Ionic Equation**: dissociate the soluble ionic compounds. Do **NOT** dissociate precipitates.

$$\overbrace{2Ag^{+}(aq) + 2NO_3^{-}(aq)}^{2AgNO_3(aq)} + \overbrace{2Na^{+}(aq) + CO_3^{2-}(aq)}^{Na_2CO_3(aq)}$$

$$\longrightarrow Ag_2CO_3(s) + \underbrace{2Na^{+}(aq) + 2NO_3^{-}(aq)}_{2NaNO_3(aq)}$$

End 10/2 G section

Precipitation Reactions and Solubility Guidelines

Write the molecular, ionic, and net ionic equations for the reaction that occurs when aqueous solutions of $AgNO_3$ and Na_2CO_3 are mixed.

End F section 10/2

4. **Net Ionic Equation**: Eliminate the spectator ions from the ionic equation.

$$2Ag^+(aq) + \cancel{2NO_3^-(aq)} + \cancel{2Na^+(aq)} + CO_3^{2-}(aq) \longrightarrow Ag_2CO_3(s) + \cancel{2Na^+(aq)} + \cancel{2NO_3^-(aq)}$$

$$\boxed{2Ag^+(aq) + CO_3^{2-}(aq) \longrightarrow Ag_2CO_3(s)}$$

HW: Precipitation Reaction For the following, Write out (a) molecular reaction (b) total ionic equation and (c) choose one of the following as the net ionic equation

Lead(II) nitrate + sodium chloride

10/3 end D section

- What is the net ionic equation for this reaction?

a. $Pb^{2+}(aq) + 2NO_3^-(aq) \longrightarrow Pb(NO_3)_2(s)$

b. $Na^{2+}(aq) + Cl^-(aq) \longrightarrow NaCl(s)$

c. $Pb^{2+}(aq) + 2Cl^-(aq) \longrightarrow PbCl_2(s)$

d. $Na^+(aq) + NO_3^-(aq) \longrightarrow NaNO_3(s)$

HW: Precipitation Reaction For the following, Write out (a) molecular reaction (b) total ionic equation and (c) choose one of the following as the net ionic equation

Lead(II) nitrate + sodium chloride

End 10/4 F, G sect
End 10/7 D section

- What is the net ionic equation for this reaction?

a. $Pb^{2+}(aq) + 2NO_3^-(aq) \longrightarrow Pb(NO_3)_2(s)$

b. $Na^{2+}(aq) + Cl^-(aq) \longrightarrow NaCl(s)$

c. $Pb^{2+}(aq) + 2Cl^-(aq) \longrightarrow PbCl_2(s)$ *

d. $Na^+(aq) + NO_3^-(aq) \longrightarrow NaNO_3(s)$

Test II ends here.

10/7 Monday was just a review day – <u>no new material covered</u> – only questions & reviewing quiz 2 answer keys (note D section had already seen the question on precipitation molecular, ionic and net ionic reaction on Thursday 10/3 their quiz date but we completed the HW in class on that HW problem for the D section on 10/7 – the answer for the precipitation problem was already posted on 10/5 before the D section did the HW problem in class for points)

Acids, Bases, and Neutralization Reactions

Acid (Arrhenius): dissociates in water to produce hydrogen ions, H^+

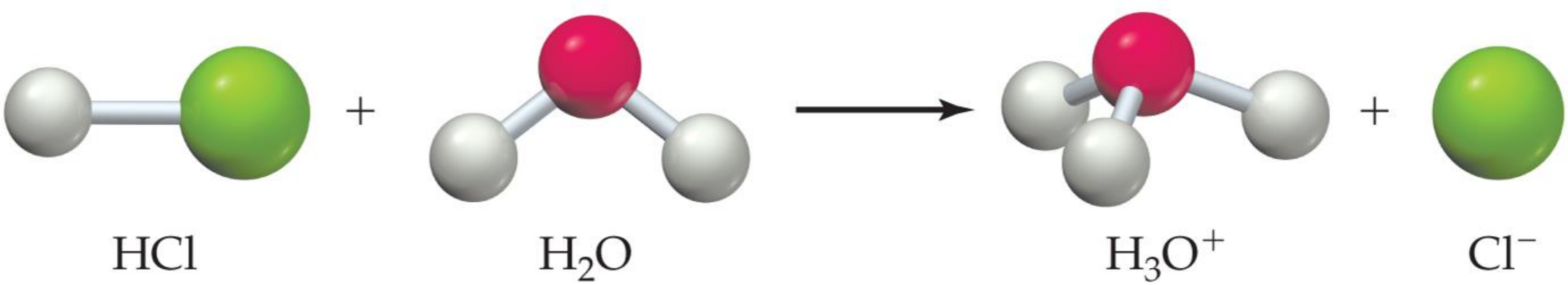
general rxn $HA(aq) \longrightarrow H^+(aq) + A^-(aq)$

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

same

In water, acids produce hydronium ions, H_3O^+:

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



Acids, Bases, and Neutralization Reactions

Base (Arrhenius): dissociates in water to produce hydroxide ions, OH^-:

general rxn $MOH(aq) \longrightarrow M^+(aq) + OH^-(aq)$

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

Ammonia is a weak base produces ammonium and hydroxide ions: (only weak base commonly seen)

$$NH_3(aq) + H_2O(aq) \longrightarrow NH_4^+(aq) + OH^-(aq)$$

Acids, Bases, and Neutralization Reactions

TABLE 4.3 Some Common Acids and Bases

Strong acid	$HClO_4$	Perchloric acid	KOH	Potassium hydroxide	Strong base
	H_2SO_4 *	Sulfuric acid	NaOH	Sodium hydroxide	
	HBr *	Hydrobromic acid	$Ba(OH)_2$	Barium hydroxide	
	HCl *	Hydrochloric acid	$Ca(OH)_2$	Calcium hydroxide	
	HNO_3 *	Nitric acid			
	H_3PO_4 *	Phosphoric acid	NH_3	Ammonia *	
	HF *	Hydrofluoric acid			
	HNO_2	Nitrous acid			
Weak acid	CH_3CO_2H *	Acetic acid			Weak base

Strong acids and **strong** bases are **strong** electrolytes.
Weak acids and **weak** bases are **weak** electrolytes.
Know name, formula, whether strong/weak acid or base for *
(most group 1A & 2A hydroxides are strong bases if soluble)

Acids – characterized by # H^+ dissociation

$H_3PO_4 \rightarrow 3 H^+ + PO_4^{-3}$ triprotic

$H_2SO_4 \rightarrow 2 H^+ + SO_4^{-2}$ diprotic

$H Br \rightarrow H^+ + Br^-$ monoprotic

$CH_3 CO_2 H \rightarrow H^+ + CH_3 CO_2^-$ monoprotic

(acetic acid – not 4 H+)

polyprotic acid = more than one hydrogen is acidic

Acids, Bases, and Neutralization Reactions

Naming Binary Acids (all halogens)

HCl **hydro**chlor**ic acid**

HBr **hydro**brom**ic acid**

HF **hydro**fluor**ic acid** **(only weak halogen acid)**

HI **hydro**iod**ic acid**

Acids, Bases, and Neutralization Reactions

These acid–base neutralization reactions are double-replacement reactions just like the precipitation reactions:

(exchange cation / anion partners)

Acid		Base		Salt		Water

$$\mathrm{HA}(aq) + \mathrm{MOH}(aq) \longrightarrow \mathrm{MA}(aq) + \mathrm{H\,O\,H}(l)$$

double replacement reaction

Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous HBr and aqueous $Ba(OH)_2$.

1. Write the chemical formulas of the products (use proper ionic rules for the salt to write neutral formula of the salt).

$$HBr(aq) + Ba(OH)_2(aq) \longrightarrow HOH + BaBr_2$$

Acid | Base | Water | Salt

Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous HBr and aqueous $Ba(OH)_2$.

2. **Molecular Equation**: Balance the equation and predict the solubility of the salt in the products. (from solubility rules table)

$$2HBr(aq) + Ba(OH)_2(aq) \longrightarrow 2H_2O(l) + \boxed{BaBr_2(aq)}$$

Use the solubility rules.

Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous HBr and aqueous $Ba(OH)_2$.

3. **Ionic Equation**: Dissociate the strong acid and the soluble ionic compounds. (solid, liquid, gas compounds, weak acid written as molecular formula-do NOT dissociate)

$$\underbrace{2H^+(aq) + 2Br^-(aq)}_{2HBr(aq)} + \underbrace{Ba^{2+}(aq) + 2OH^-(aq)}_{Ba(OH)_2(aq)}$$

$$\longrightarrow 2H_2O(l) + \underbrace{Ba^{2+}(aq) + 2Br^-(aq)}_{BaBr_2(aq)}$$

Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous HBr and aqueous $Ba(OH)_2$.

4. **Net Ionic Equation**: Eliminate the spectator ions from the ionic equation.

$$2H^+(aq) + 2Br^-(aq) + Ba^{2+}(aq) + 2OH^-(aq)$$

$$\longrightarrow 2H_2O(l) + Ba^{2+}(aq) + 2Br^{1-}(aq)$$

$$2H^+(aq) + 2OH^-(aq) \longrightarrow 2H_2O(l)$$

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

Net ionic equation for acid/base neutralization is almost always this (unless weak undissociated acid/base)

Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous NaOH and aqueous HF.

1. Write the chemical formulas of the products (use proper ionic rules for the salt to write neutral ionic formulas). (exchange cation anion partners)

$$HF(aq) + NaOH(aq) \longrightarrow HOH + NaF$$

Acid Base Water Salt

Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous NaOH and aqueous HF.

2. **Molecular Equation**: Balance the equation and predict the solubility of the salt in the products.

$$HF(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaF(aq)$$

use the solubility rules.

Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous NaOH and aqueous HF.

3. **Ionic Equation**: Dissociate the soluble ionic compounds. (HF is a weak acid so does not dissociate)

$$\mathrm{HF}(aq) + \overbrace{\mathrm{Na}^{+}(aq) + \mathrm{OH}^{-}(aq)}^{\mathrm{NaOH}(aq)} \longrightarrow \mathrm{H_2O}(l) + \underbrace{\mathrm{Na}^{+}(aq) + \mathrm{F}^{-}(aq)}_{\mathrm{NaF}(aq)}$$

Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous NaOH and aqueous HF.

4. **Net Ionic Equation**: Eliminate the spectator ions from the ionic equation.

$$HF(aq) + \cancel{Na^{+}(aq)} + OH^{-}(aq) \longrightarrow H_2O(l) + \cancel{Na^{+}(aq)} + F^{-}(aq)$$

$$\boxed{HF(aq) + OH^{-}(aq) \longrightarrow H_2O(l) + F^{-}(aq)}$$

HW: Acids, Bases, and Neutralization Reactions

Write the molecular, ionic, and net ionic equations for the reaction of aqueous $Ca(OH)_2$ and aqueous HI.

molecular equation:

$Ca(OH)_2$ + HI → ____________ + ______________

ionic equation

net ionic equation

Solution Stoichiometry(already know how to do combining 2 things you know – stoichiometry & Molarity)

What volume of 0.250 M H_2SO_4 is needed to react with 50.0 mL of 0.100 M NaOH? 10.0 mL H_2SO_4

$$H_2SO_4(aq) + 2NaOH(aq) \longrightarrow Na_2SO_4(aq) + 2H_2O(l)$$

End 10/11 F F section

What volume of 0.250 M H Cl is needed to react with 50.0 mL of 0.100 M NaOH? 20.0 mL HCl

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

End 10/10/19 D, G section

Solution Stoichiometry (short cut version)

What volume of 0.250 M H Cl is needed to react with 50.0 mL of 0.100 M NaOH?

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

Short cut: use $M_{acid}V_{acid} = M_{base}V_{base}$

only works for 1:1 acid/base reaction

$$(0.250\text{ M HCl}) * (V_{HCl}) = (0.100\text{ M NaOH}) * (50.0\text{ mL NaOH})$$

$$V_{HCl} = \frac{(0.100\text{ M NaOH}) * (50.0\text{ mL NaOH})}{0.250\text{ M H Cl}} = 20.0\text{ mL HCl}$$

HW: Solution Stoichiometry

What volume of 2.25 M $Ca(OH)_2$ is needed to react with 25.0 mL of 0.100 M HCl ?

$$2\,HCl(aq) + Ca(OH)_2(aq) \longrightarrow CaCl_2(aq) + 2HOH(l)$$

HW: Solution Stoichiometry

What volume of 2.25 M Ca $(OH)_2$ is needed to react with 25.0 mL of 0.100 M H Cl ?

$$2\ HCl(aq) + Ca(OH)_{2}(aq) \longrightarrow CaCl_2(aq) + 2HOH(l)$$

$$\frac{25.0\ \text{mL}}{\text{HCl soln}} \times \frac{0.100\ \text{moles H Cl}}{1000\ \text{mL HCl soln}} \times \frac{1\ \text{mol}\ Ca(OH)_2}{2\ \text{mol H Cl}} \times \frac{1000\ \text{mL}\ Ca(OH)_2}{2.25\ \text{moles}\ Ca(OH)_2}$$

= 0.556 mL $Ca(OH)_2$ solution

Measuring the Concentration of a Solution: Titration

Titration: A procedure for determining the concentration using a second solution with known concentration

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

Short cut: use $M_{acid}V_{acid} = M_{base}V_{base}$

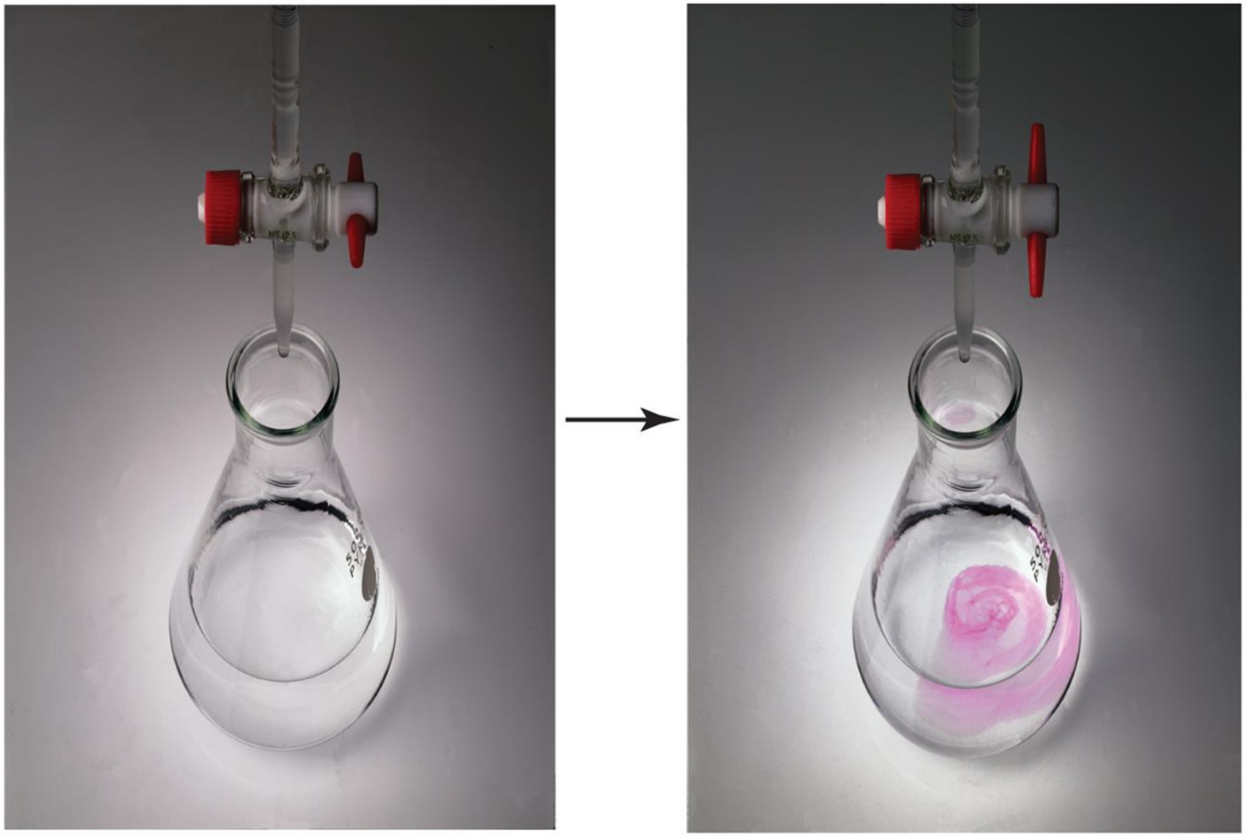
How can you tell when the reaction is complete?

Equivalence point – point where you have the same amount H^+ and OH^-

Indicator: Used to mark the equivalence point

changes color at the equivalence point

Measuring the Concentration of a Solution: Titration

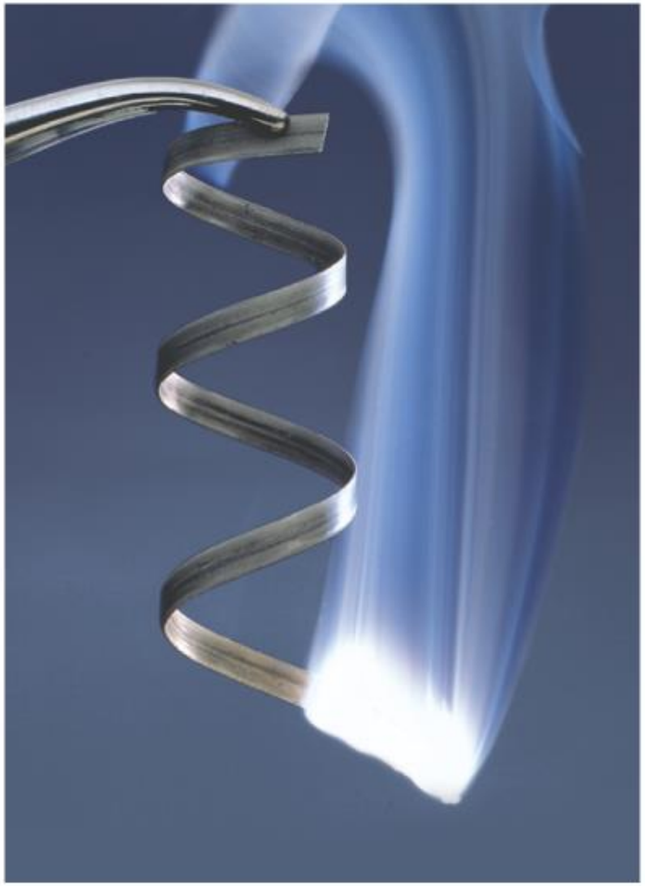


A measured volume of acid solution is placed in a flask, and phenolphthalein indicator is added.

Base solution of known concentration is added from a buret until the indicator changes **color**. Reading the volume of base from the buret allows calculation of the acid concentration.

Oxidation–Reduction (Redox) Reactions

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



Oxidation–Reduction (Redox) Reactions

$$2P(s) + 3Br_2(l) \longrightarrow 2PBr_3(l)$$

Oxidation–Reduction (Redox) Reactions

zero minus electrons +3

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

Rusting of iron:
an **oxidation** of Fe

Oxidation → lose electrons

$$A^{2-} \longrightarrow A^{-} + \text{electron}$$

$$A^{-} \longrightarrow A + \text{electron}$$

$$A \longrightarrow A^{+} + \text{electron}$$

$$A^{+} \longrightarrow A^{2+} + \text{electron}$$

← *Reduction* gain electrons

Reactant A might be:
a neutral atom,
a monatomic ion,
a polyatomic ion,
or a molecule.

$$2Fe_2O_3(s) + 3C(s) \longrightarrow 4Fe(s) + 3CO_2(g)$$

+3 add electrons zero

Manufacture of iron:
a **reduction** of Fe_2O_3

Oxidation–Reduction (Redox) Reactions

Oxidation: The loss of one or more electrons by a substance, whether element, compound, or ion

Reduction: The gain of one or more electrons by a substance (memorize as add negative – reduced ox number)

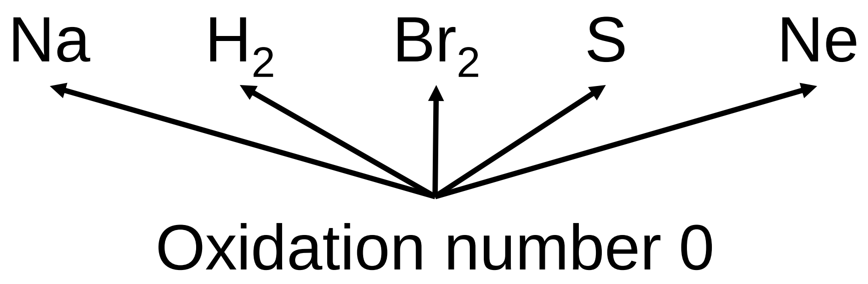
Oxidation–Reduction (Redox) Reaction: Any process in which electrons are transferred from one substance to another

Oxidation–Reduction (Redox) Reactions

Oxidation Number (State): A value that indicates whether an atom is neutral, electron-rich, or electron-poor (These rules are on the memorize list on departmental syllabus.)

Rules for Assigning Oxidation Numbers

1. An atom in its elemental state has an oxidation number of 0 (zero).



Oxidation–Reduction (Redox) Reactions

2. An atom in a monatomic ion has an oxidation number identical to its charge.

   (group 1A to 3A: group #) or

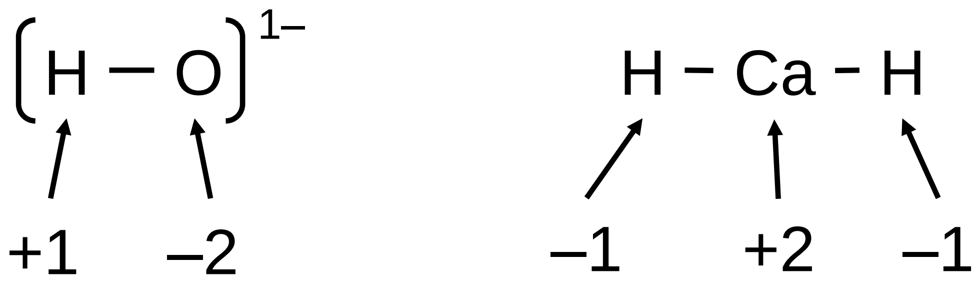
   (group 7A to 5A: group # - 8)

1A	2A	3A	7A	6A	5A
Na^+	Ca^{2+}	Al^{3+}	Cl^-	O^{2-}	N^{3-}
+1	+2	+3	–1	–2	–3

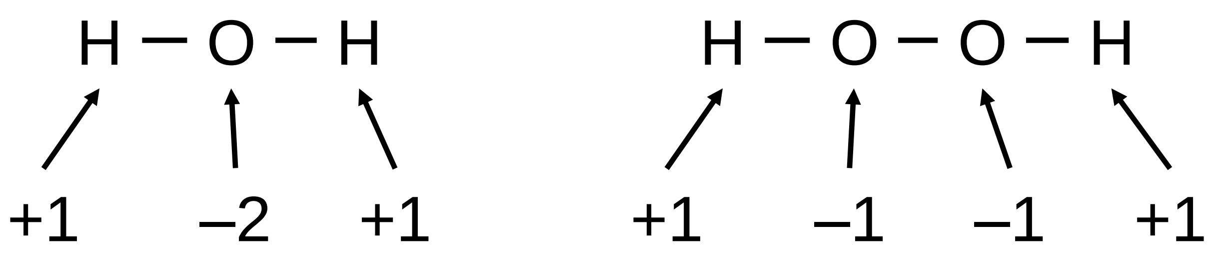
Oxidation–Reduction (Redox) Reactions

3. An atom in a polyatomic ion or in a molecular compound usually has the same oxidation number it would have if it were a monatomic ion.

   a) Hydrogen can be either +1 or –1.



   b) Oxygen *usually* has an oxidation number of –2.



Oxidation–Reduction (Redox) Reactions

3. c) Halogens *usually* have an oxidation number of –1.

Oxidation–Reduction (Redox) Reactions

4. The sum of the oxidation numbers is 0 for a neutral compound and is equal to the net charge for a polyatomic ion.

End 10/16 W D section

H_2SO_4

+1 ? –2

$2(+1) + (?) + 4(-2) = 0$ (net charge)

$S = ? = +6$

$Cr_2O_7^{2-}$

? –2

$2(?) + 7(-2) = -2$ (net charge)

$Cr = ? = +6$

What is the oxidation state of the following.

a. Na element zero

b. H_2 element zero

c. Na Cl Na+1(gp 1A) 7-8 = -1 (gp 7A)

d. HNO_3 H +1(gp 1A), O -2 (gp 6A)

ox state N = variable N, calculate

zero = +1 + N + 3*(-2)

algebra N = +6 - 1 = 5

HW: What is the oxidation state of the following.

a. Cl_2

b. Fe

End 10/16 F section, G section

c. $AlCl_3$

d. PO_4^{-3}

Identifying Redox Reactions

Oxidizing Agent (is reduced)

Causes oxidation

Gains one or more electrons

Undergoes reduction

Oxidation number of atom decreases.

(becomes more negative)

Reducing Agent (is oxidized)

Causes reduction

Loses one or more electrons

Undergoes oxidation

Oxidation number of atom increases.

(becomes more positive)

Identifying Redox Reactions

this chapter - only responsible for recognizing oxidation/reduction & oxidizing agent & reducing agent – NOT for balancing redox equations (chapter 18)

Oxidizing Agent

Reduction

6(+ 2 e)

0 –2

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

0 +3

4(- 3 e)

Oxidation

Reducing Agent

Identifying Redox Reactions

Oxidizing Agent

Reduction

+3 0

$$2Fe_2O_3(s) + 3C(s) \longrightarrow 4Fe(s) + 3CO_2(g)$$

0 +4

Oxidation

Reducing Agent

HW: Oxidation or Reduction

(reducing agent / oxidizing agent)

(a) Assign ox states of Zn & H

(b) assign oxidation or reduction to brackets

$$Zn + 2\,HCl \rightarrow ZnCl_2 + H_2$$

HW: Oxidation or Reduction

(reducing agent / oxidizing agent)

(a) Assign ox states of Zn & H

(b) assign oxidation or reduction to brackets

+ 1e per H reduction (H is oxidizing agent)

zero +2

$$Zn + 2\,HCl \rightarrow ZnCl_2 + H_2$$

+1 zero

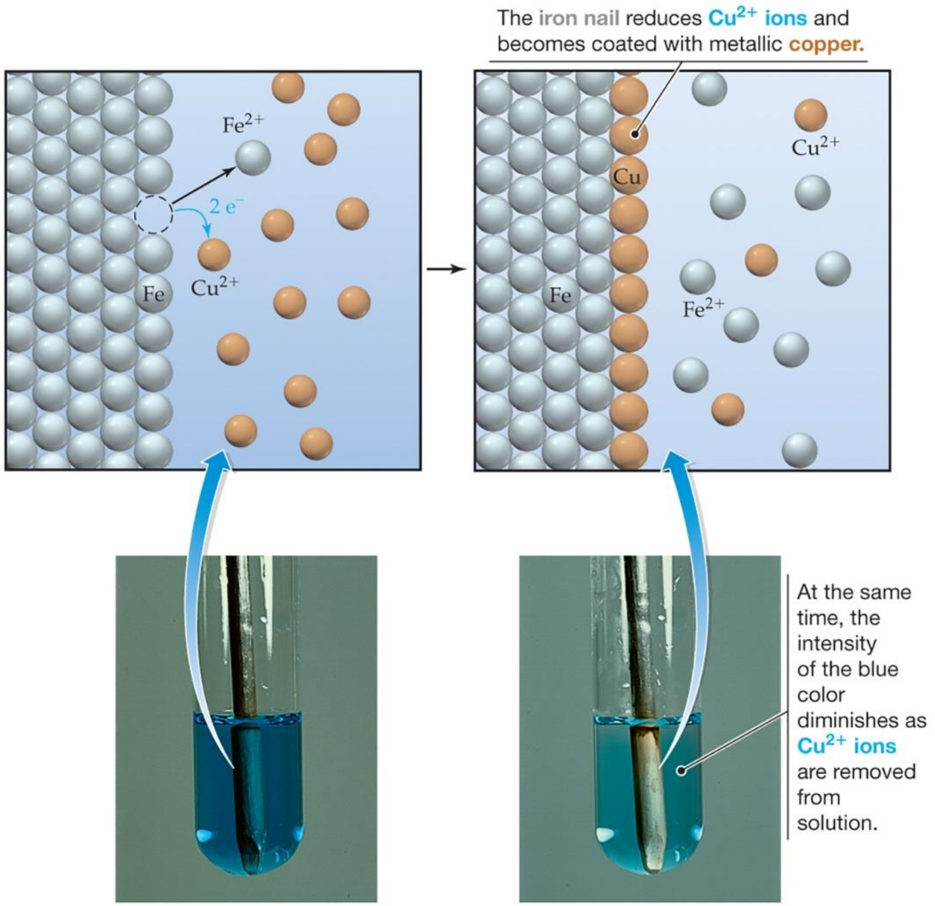
- 2e per Zn oxidation (Zn is reducing agent)

10/17 R - D section

The Activity Series of the Elements

Iron loses electron to copper

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



The Activity Series of the Elements

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

TABLE 4.5 A Partial Activity Series of the Elements

Oxidation Reaction		
Strongly reducing	These elements react rapidly with aqueous H^+ ions (acid) or with liquid H_2O to release H_2 gas.	$Li \longrightarrow Li^+ + e^-$ $K \longrightarrow K^+ + e^-$ $Ba \longrightarrow Ba^{2+} + 2e^-$ $Ca \longrightarrow Ca^{2+} + 2e^-$ $Na \longrightarrow Na^+ + e^-$
	These elements react with aqueous H^+ ions or with steam to release H_2 gas.	$Mg \longrightarrow Mg^{2+} + 2e^-$ $Al \longrightarrow Al^{3+} + 3e^-$ $Mn \longrightarrow Mn^{2+} + 2e^-$ $Zn \longrightarrow Zn^{2+} + 2e^-$ $Cr \longrightarrow Cr^{3+} + 3e^-$ $Fe \longrightarrow Fe^{2+} + 2e^-$
	These elements react with aqueous H^+ ions to release H_2 gas.	$Co \longrightarrow Co^{2+} + 2e^-$ $Ni \longrightarrow Ni^{2+} + 2e^-$ $Sn \longrightarrow Sn^{2+} + 2e^-$
		$H_2 \longrightarrow 2H^+ + 2e^-$
Weakly reducing	These elements do not react with aqueous H^+ ions to release H_2.	$Cu \longrightarrow Cu^{2+} + 2e^-$ $Ag \longrightarrow Ag^+ + e^-$ $Hg \longrightarrow Hg^{2+} + 2e^-$ $Pt \longrightarrow Pt^{2+} + 2e^-$ $Au \longrightarrow Au^{3+} + 3e^-$

Elements that are higher up in the table (reaction goes →) are more likely to be oxidized (lose e).

Thus, any element higher in the activity series will reduce (reducing agent) (**<u>lose e to</u>**) the ion of any element lower in the activity series.

The Activity Series of the Elements

$Cu(s) + 2Ag^{+}(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$ (a)

$2Ag(s) + Cu^{2+}(aq) \longrightarrow 2Ag^{+}(aq) + Cu(s)$ (b)

Which one of these reactions will occur?

TABLE 4.5 A Partial Activity Series of the Elements

Oxidation Reaction		
Strongly reducing	These elements react rapidly with aqueous H^+ ions (acid) or with liquid H_2O to release H_2 gas.	$Li \longrightarrow Li^+ + e^-$
		$K \longrightarrow K^+ + e^-$
		$Ba \longrightarrow Ba^{2+} + 2e^-$
		$Ca \longrightarrow Ca^{2+} + 2e^-$
		$Na \longrightarrow Na^+ + e^-$
	These elements react with aqueous H^+ ions or with steam to release H_2 gas.	$Mg \longrightarrow Mg^{2+} + 2e^-$
		$Al \longrightarrow Al^{3+} + 3e^-$
		$Mn \longrightarrow Mn^{2+} + 2e^-$
		$Zn \longrightarrow Zn^{2+} + 2e^-$
		$Cr \longrightarrow Cr^{3+} + 3e^-$
		$Fe \longrightarrow Fe^{2+} + 2e^-$
	These elements react with aqueous H^+ ions to release H_2 gas.	$Co \longrightarrow Co^{2+} + 2e^-$
		$Ni \longrightarrow Ni^{2+} + 2e^-$
		$Sn \longrightarrow Sn^{2+} + 2e^-$
		$H_2 \longrightarrow 2H^+ + 2e^-$
	These elements do not react with aqueous H^+ ions to release H_2.	$Cu \longrightarrow Cu^{2+} + 2e^-$
		$Ag \longrightarrow Ag^+ + e^-$
		$Hg \longrightarrow Hg^{2+} + 2e^-$
		$Pt \longrightarrow Pt^{2+} + 2e^-$
Weakly reducing		$Au \longrightarrow Au^{3+} + 3e^-$

Cu is above Ag in activity series

Cu → Ag ←

The Activity Series of the Elements

$Cu(s) + 2Ag^{+}(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$ (a) *

$2Ag(s) + Cu^{2+}(aq) \longrightarrow 2Ag^{+}(aq) + Cu(s)$ (b)

Which one of these reactions will occur?

TABLE 4.5 A Partial Activity Series of the Elements

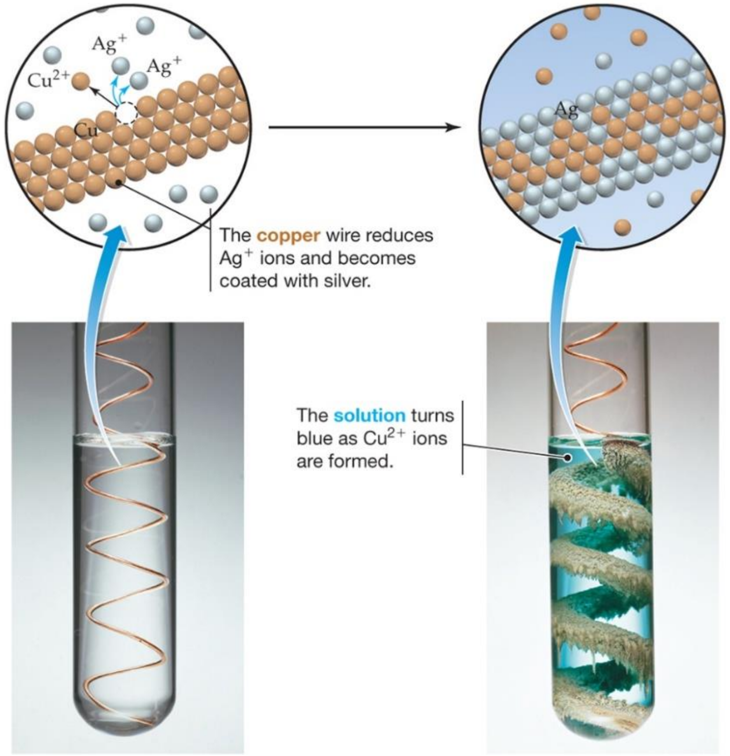
		Oxidation Reaction
Strongly reducing	These elements react rapidly with aqueous H^+ ions (acid) or with liquid H_2O to release H_2 gas.	$Li \longrightarrow Li^+ + e^-$
		$K \longrightarrow K^+ + e^-$
		$Ba \longrightarrow Ba^{2+} + 2e^-$
		$Ca \longrightarrow Ca^{2+} + 2e^-$
		$Na \longrightarrow Na^+ + e^-$
	These elements react with aqueous H^+ ions or with steam to release H_2 gas.	$Mg \longrightarrow Mg^{2+} + 2e^-$
		$Al \longrightarrow Al^{3+} + 3e^-$
		$Mn \longrightarrow Mn^{2+} + 2e^-$
		$Zn \longrightarrow Zn^{2+} + 2e^-$
		$Cr \longrightarrow Cr^{3+} + 3e^-$
		$Fe \longrightarrow Fe^{2+} + 2e^-$
	These elements react with aqueous H^+ ions to release H_2 gas.	$Co \longrightarrow Co^{2+} + 2e^-$
		$Ni \longrightarrow Ni^{2+} + 2e^-$
		$Sn \longrightarrow Sn^{2+} + 2e^-$
		$H_2 \longrightarrow 2H^+ + 2e^-$
	These elements do not react with aqueous H^+ ions to release H_2.	$Cu \longrightarrow Cu^{2+} + 2e^-$
		$Ag \longrightarrow Ag^+ + e^-$
		$Hg \longrightarrow Hg^{2+} + 2e^-$
		$Pt \longrightarrow Pt^{2+} + 2e^-$
Weakly reducing		$Au \longrightarrow Au^{3+} + 3e^-$

Cu is above Ag in activity series

Cu → Ag ←

The Activity Series of the Elements

$$Cu(s) + 2Ag^{+}(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$$

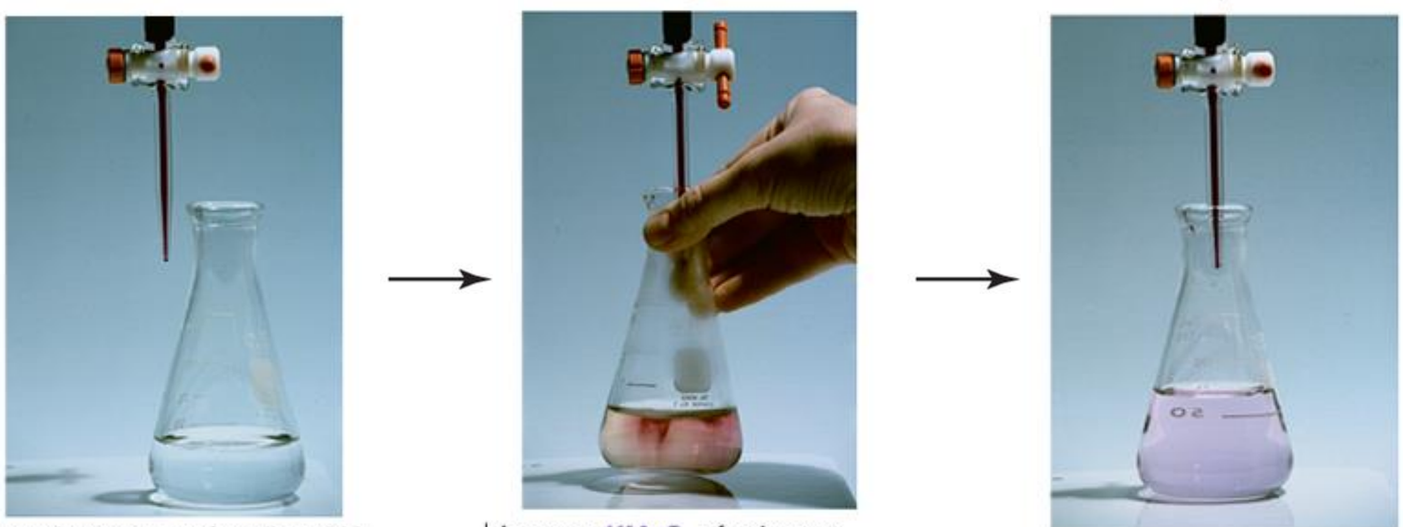


copper is above silver in activity series chart

(copper loses electron and silver gets electron)

Redox Titrations

$$5H_2C_2O_4(aq) + 2MnO_4^-(aq) + 6H^+(aq) \longrightarrow 10CO_2(g) + 2Mn^{2+}(aq) + 8H_2O(l)$$



A precise amount of oxalic acid is weighed and dissolved in aqueous H_2SO_4.

Aqueous $KMnO_4$ of unknown concentration is added from a buret until …

… the purple color persists, indicating that all of the oxalic acid has reacted.

Redox Titrations (probably will not have on exam/quiz)

A solution is prepared with 0.2585 g of oxalic acid, $H_2C_2O_4$. 22.35 mL of an unknown solution of potassium permanganate is needed to titrate the solution. What is the concentration of the potassium permanganate solution? (already know how to do this: if given balanced reaction, just stoichiometry & molarity problem)

$$5H_2C_2O_4(aq) + 2MnO_4^-(aq) + 6H^+(aq) \longrightarrow 10CO_2(g) + 2Mn^{2+}(aq) + 8H_2O(l)$$

Mass of $H_2C_2O_4$ → Moles of $H_2C_2O_4$ → Moles of $KMnO_4$ → Molarity of $KMnO_4$

Molar mass of $H_2C_2O_4$ | Mole ratio | Molarity of $KMnO_4$

Redox Titrations

$$5H_2C_2O_4(aq) + 2MnO_4^-(aq) + 6H^+(aq) \longrightarrow 10CO_2(g) + 2Mn^{2+}(aq) + 8H_2O(l)$$

Moles of $H_2C_2O_4$ available:

$$0.2585 \text{ g } H_2C_2O_4 \times \frac{1 \text{ mol}}{90.04 \text{ g}} = 0.002871 \text{ mol } H_2C_2O_4$$

Moles of $KMnO_4$ reacted:

$$0.002871 \text{ mol } H_2C_2O_4 \times \frac{2 \text{ mol } KMnO_4}{5 \text{ mol } H_2C_2O_4} = 0.001148 \text{ mol } KMnO_4$$

Redox Titrations

$$5H_2C_2O_4(aq) + 2MnO_4^-(aq) + 6H^+(aq) \longrightarrow 10CO_2(g) + 2Mn^{2+}(aq) + 8H_2O(l)$$

Concentration of $KMnO_4$ solution:

$$\frac{0.001148 \text{ mol } KMnO_4}{22.35 \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = \boxed{0.05136 \text{ M } KMnO_4}$$