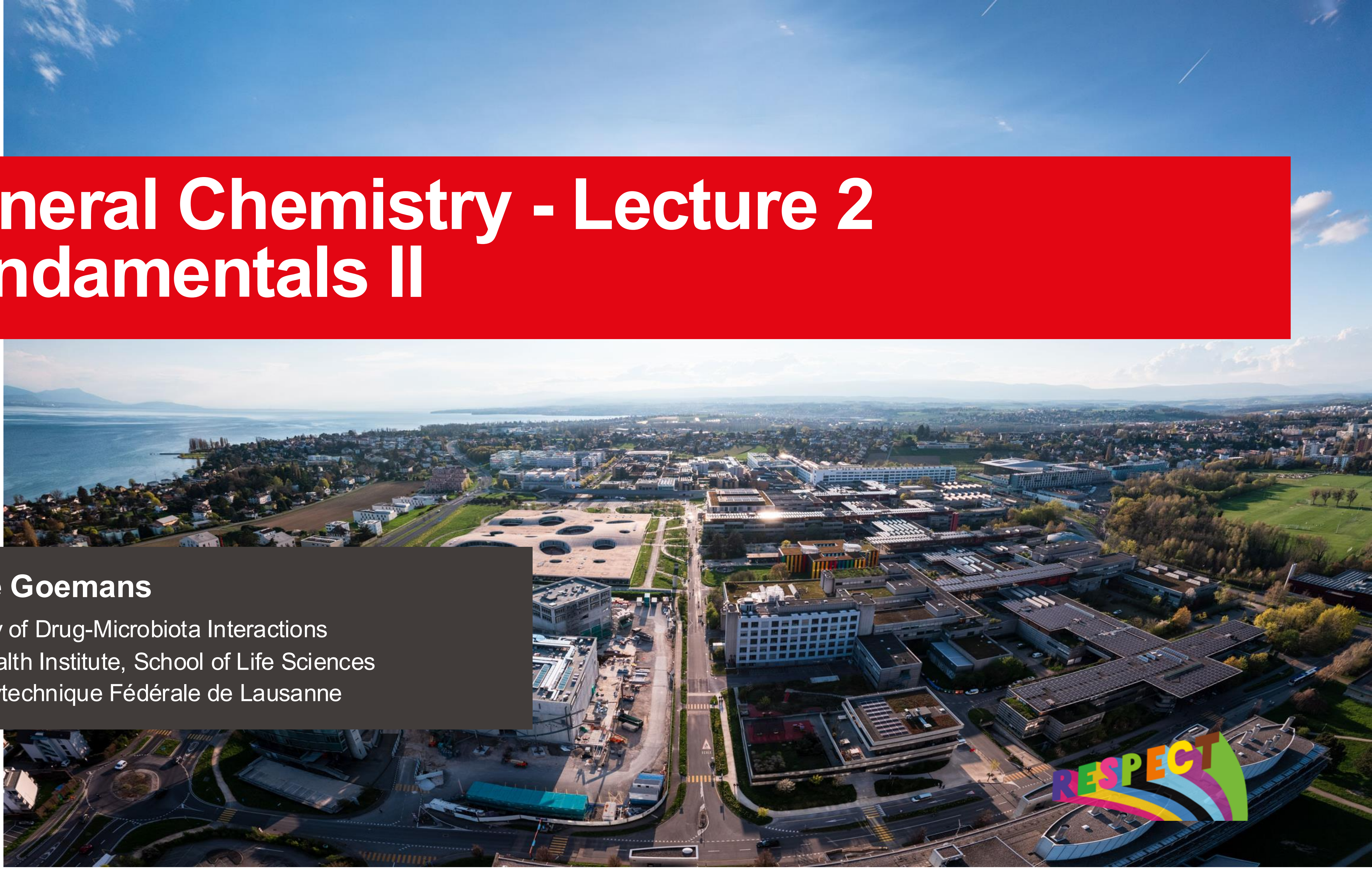

General Chemistry - Lecture 2
Fundamentals II

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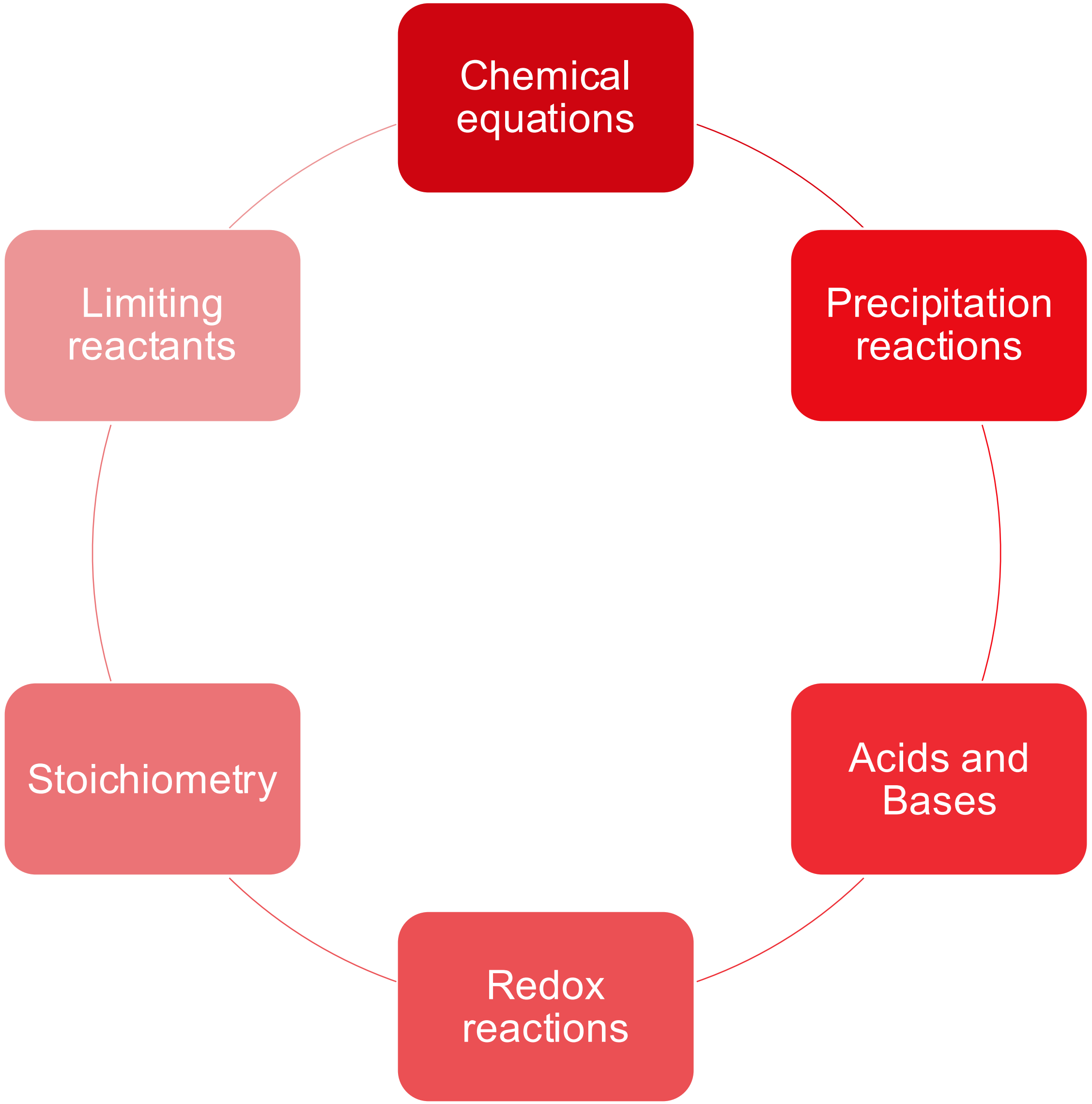
Plan



Chemical equations

Precipitation reactions

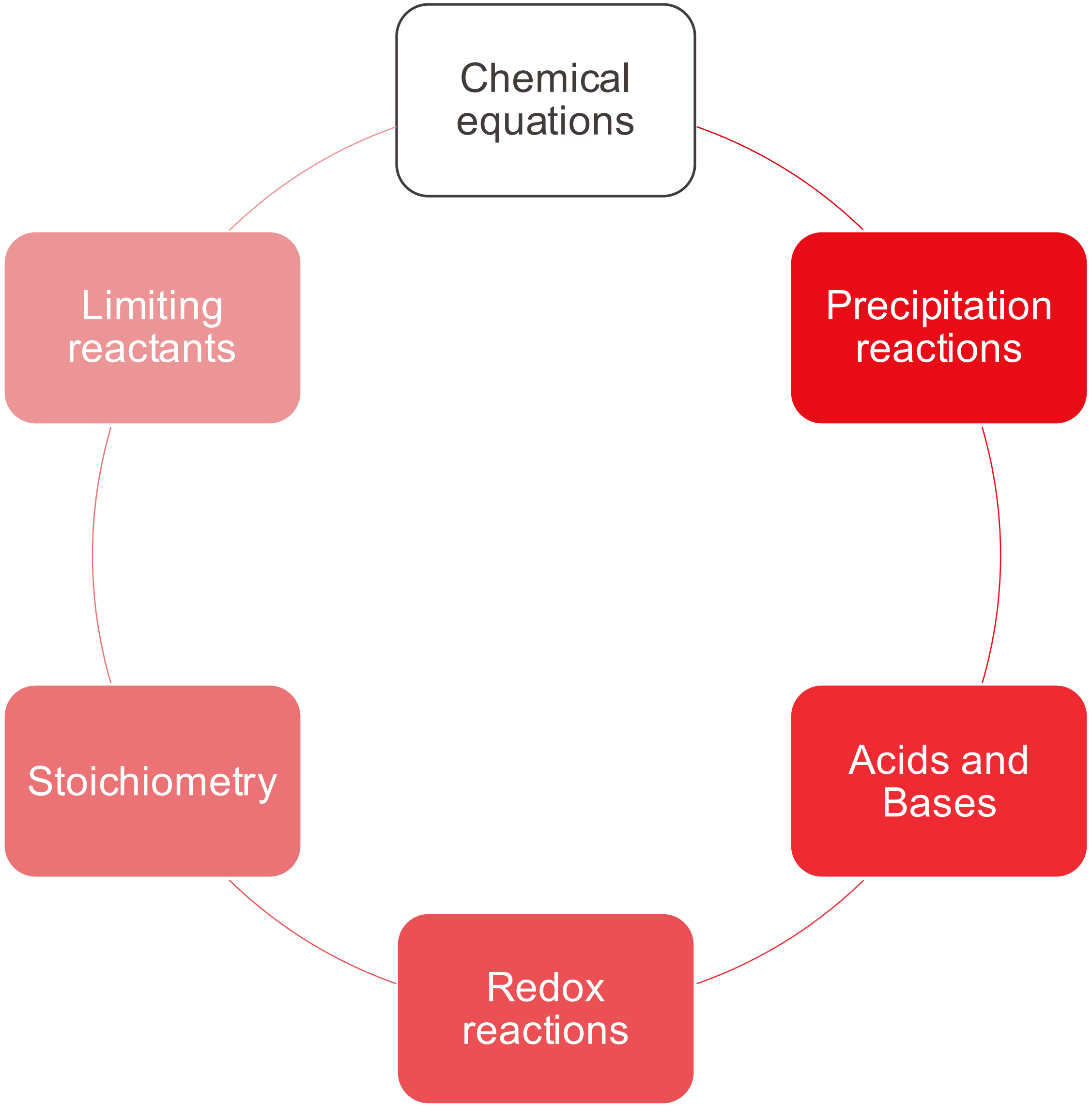
Acids and Bases

Redox reactions

Stoichiometry

Limiting reactants

Plan



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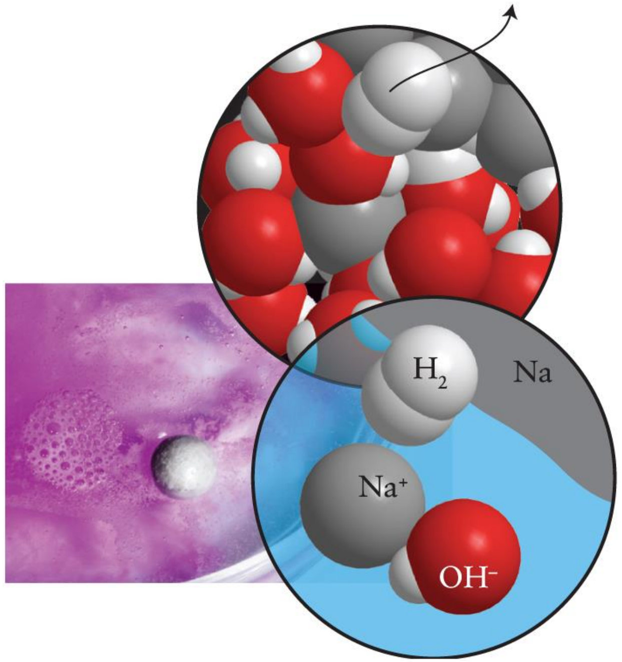
Different ways to describe a chemical reaction

- When we drop a small lump of sodium metal into a container of water, hydrogen gas forms rapidly and sodium hydroxide is produced in solution.
- Words: Sodium + water → sodium hydroxide + hydrogen
- Chemical formula: $Na + H_2O \rightarrow NaOH + H_2$
- The skeletal equation is a **qualitative summary** of a chemical reaction that is not yet balanced.



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Chemical reactions

- Chemical reactions are symbolized by an arrow (→):

reactants → products

- This is known as a **skeletal equation** (“bare bones” equation). It is a *qualitative equation*.
- To *quantitatively* show chemical reactions, the atoms must obey the **law of conservation of mass**. The atoms are not created or destroyed but may be rearranged into different combinations.
- When the same number of each type of atom is present on each side of the equation arrow, the **chemical equation** is *balanced*.

Chemical reactions

$$2\ Na + 2\ H_2O \rightarrow 2\ NaOH + H_2$$

- A **subscript** gives the number of each type of atom.
- A **stoichiometric coefficient** shows the number of each atom, ion, or compound. Since there are two moles of Na, two moles of O, and four moles of H on each side of the equation, it is *balanced*.
- Since the equation is balanced, it adheres to the **law of conservation of mass**.
- A complete chemical equation also typically shows the **physical state** of each reactant and product by using a state symbol:

$$2\ Na\ (\mathbf{s}) + 2\ H_2O\ (\mathbf{l}) \rightarrow 2\ NaOH\ (\mathbf{aq}) + H_2\ (\mathbf{g})$$

Chemical reactions

$2\ Na\ (\mathbf{s}) + 2\ H_2O\ (\mathbf{l}) \rightarrow 2\ NaOH\ (\mathbf{aq}) + H_2\ (\mathbf{g})$

Using Avogadro's principle, we conclude: 2 moles of Na atoms react with 2 moles of H_2O molecules, producing 2 moles of NaOH formula units and 1 mole of H_2 molecules.

The **stoichiometric coefficients** tell us the ***relative number*** of moles of each substance that reacts or is produced in the reaction.

Chemical reactions – additional information

- The Greek letter Δ (delta) over the arrow indicates heat.
  - Converting limestone into quicklime requires 800°C: $CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$
- A catalyst is a substance that increases the rate of a reaction but is not itself consumed in the reaction.
  - Here vanadium (V) oxide is a catalyst during the production of sulfuric acid.

$$2\,SO_2(g) + O_2(g) \xrightarrow{V_2O_5} 2\,SO_3(g)$$

Balancing chemical reactions

- Stoichiometric coefficients balance equations
- We start with a skeletal equation and count the atoms on each side.

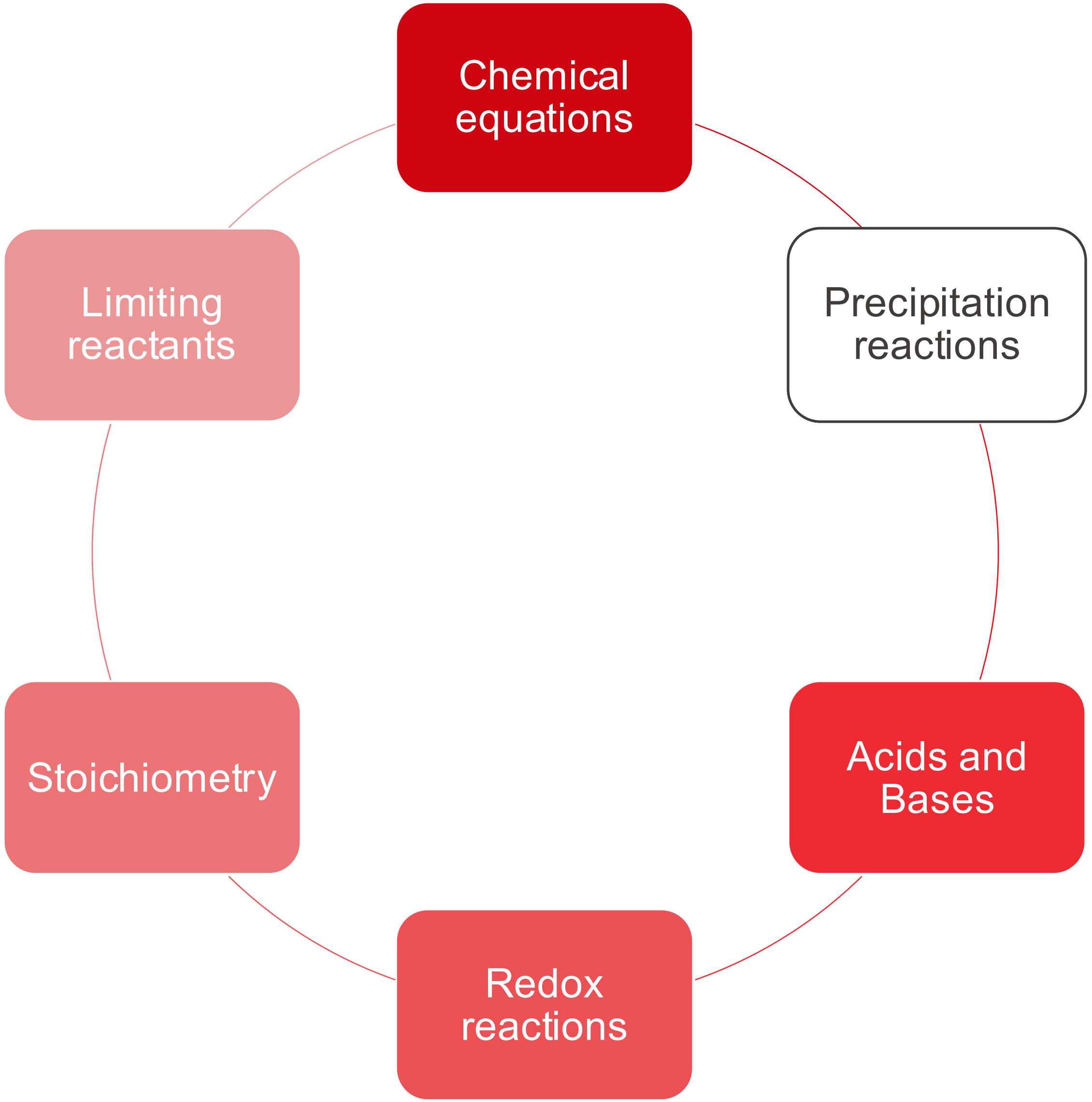
$$H_2 + O_2 \rightarrow H_2O$$

- On the reactant side of the equation, there are 2 hydrogens and 2 oxygens. However, on the product side, there is only 1 oxygen while there are 2 hydrogens. This violates the law of conservations of mass.
- We place a 2 in front of both the hydrogen and water, then recount atoms.

$$2\ H_2 + O_2 \rightarrow 2\ H_2O$$

- Now, on the reactant side of the equation, there are 4 hydrogens and 2 oxygens, and the product side has the same. Therefore, this adheres to the law of conservation of mass.
- Never add or remove **subscript** when balancing equations!

Plan



Chemical equations

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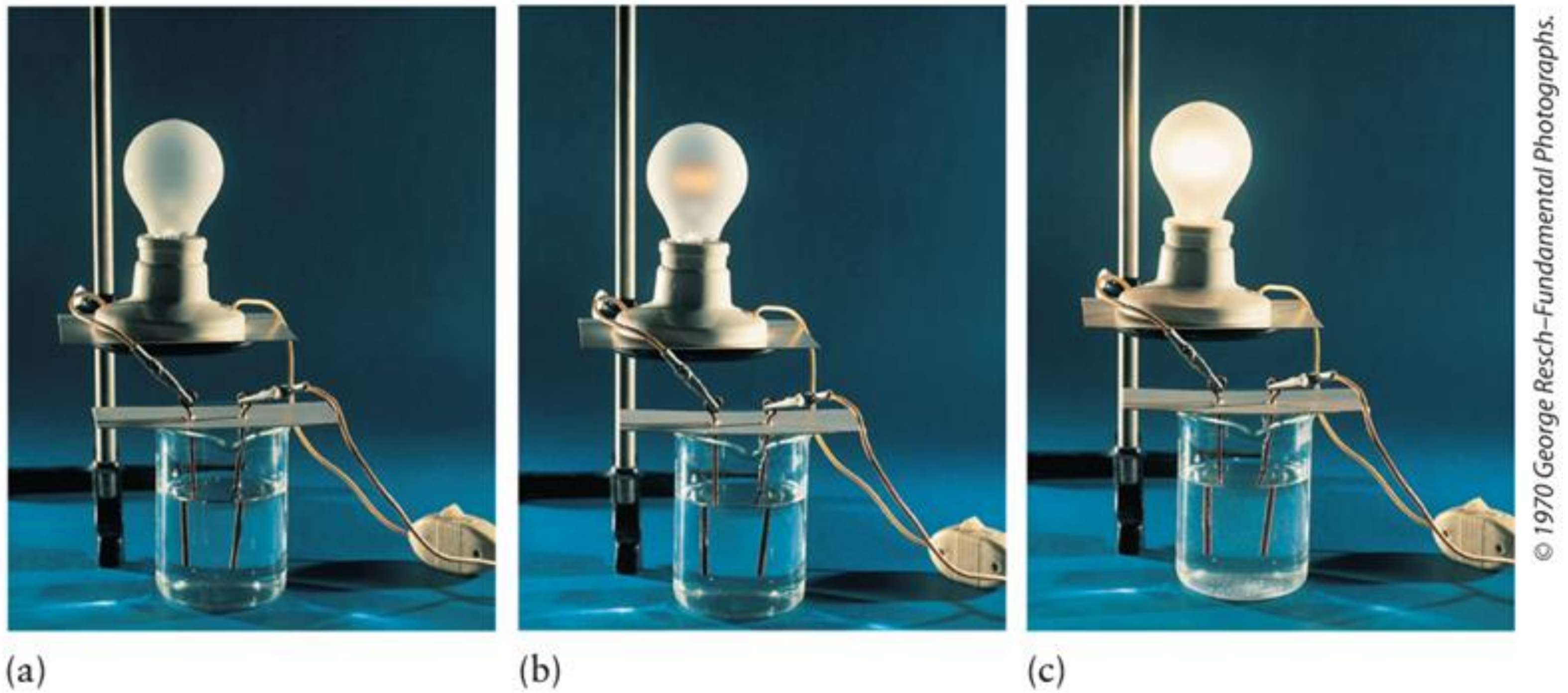
Electrolyte solutions

When dissolved in water, some compounds conduct electricity—some strongly and some not at all.

nonelectrolyte weak electrolyte **strong electrolyte**



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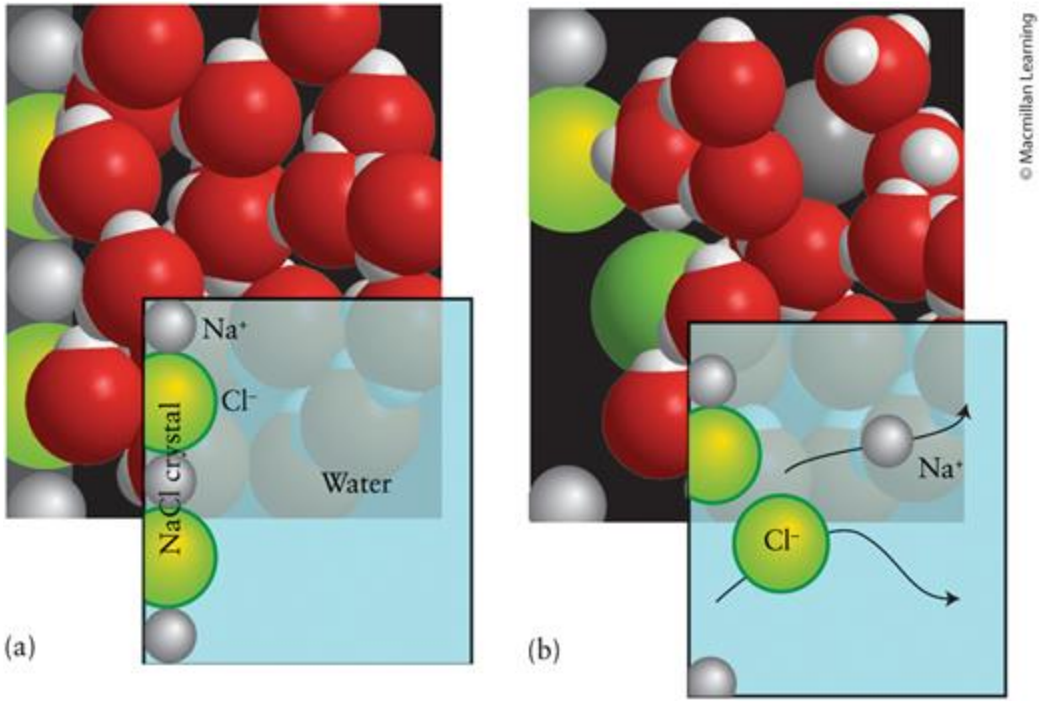
Electrolyte solutions

- Electrolytes conduct electricity by migration of ions.
- Only solutions with ions can conduct.
- When ionic solids dissolve, ions become free to migrate.

Non-migrating ions



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Ions free to migrate

Strong or weak electrolytes

- **Strong electrolytes** are present almost entirely as ions in solution
- Three types:
  - Strong acids (See *later*)
  - Strong bases (See *later*)
  - Soluble ionic compounds
- **Weak electrolytes** ionize incompletely in solution.
  - Most of the molecules remain intact.
  - For example: Acetic acid dissolves mainly as molecules with only a small fraction of H^+ and acetate ions present.

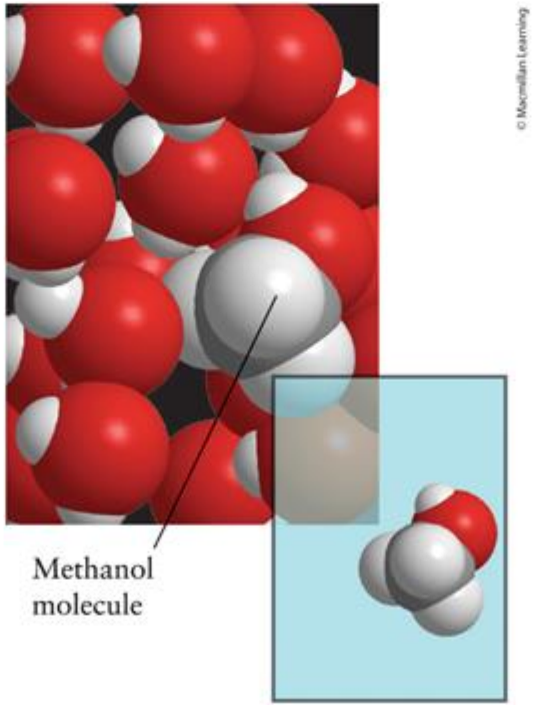
Non-electrolyte solutions

Non-electrolytes

- Do not form ions in solution
- May be soluble or not.
- If soluble, dissolves as intact molecules.
- Does NOT dissociate

Non-electrolyte solutions

- Solutions do NOT conduct electricity.



Dissolving vs. dissociating

- **Sugar** dissolves in water as molecules; it does not dissociate into ions.
- **Oxygen** dissolves in your blood as molecules; it does not ionize.
- **NaCl** dissociates in water to form separate ions and an electrolyte solution.
- **NaOH** dissociates in water to form separate ions and an electrolyte solution.
- **Acetic acid** dissolves in water mostly as molecules and dissociates to form a very few H^+ and acetate ions.

Forming a new solid: precipitation reactions

Adding a sodium chloride solution to a solution of silver nitrate results in the immediate formation of a new solid, silver chloride.

- How does this happen?

- When you mix solutions of silver nitrate and sodium chloride, a precipitate forms.

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(\mathbf{s}) + NaNO_3(aq)$$

- In a precipitation reaction, **(aq)** indicates a substance dissolved in water and **(s)** indicates a solid that has precipitated.

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Ionic and net ionic equations

$$AgNO_3\ (a) + NaCl\ (aq) \rightarrow AgCl\ (s) + NaNO_3\ (aq)$$

Rewrite the **molecular equation** showing the **dissociated** ions to give the **complete ionic equation:**

$$Ag^+\ (aq) + NO_3^-\ (aq) + Na^+\ (aq) + Cl^-(aq) \rightarrow AgCl\ (s) + Na^+\ (aq) + NO_3^-\ (aq)$$

We notice the same ions on each side of the equation, which means they play **no role** in our chemical reaction. These are called **spectator ions**.

Rewrite the **complete ionic equation without the spectators** to give the **net ionic equation**.

$$Ag^+\ (aq) + Cl^-\ (aq) \rightarrow AgCl\ (s)$$

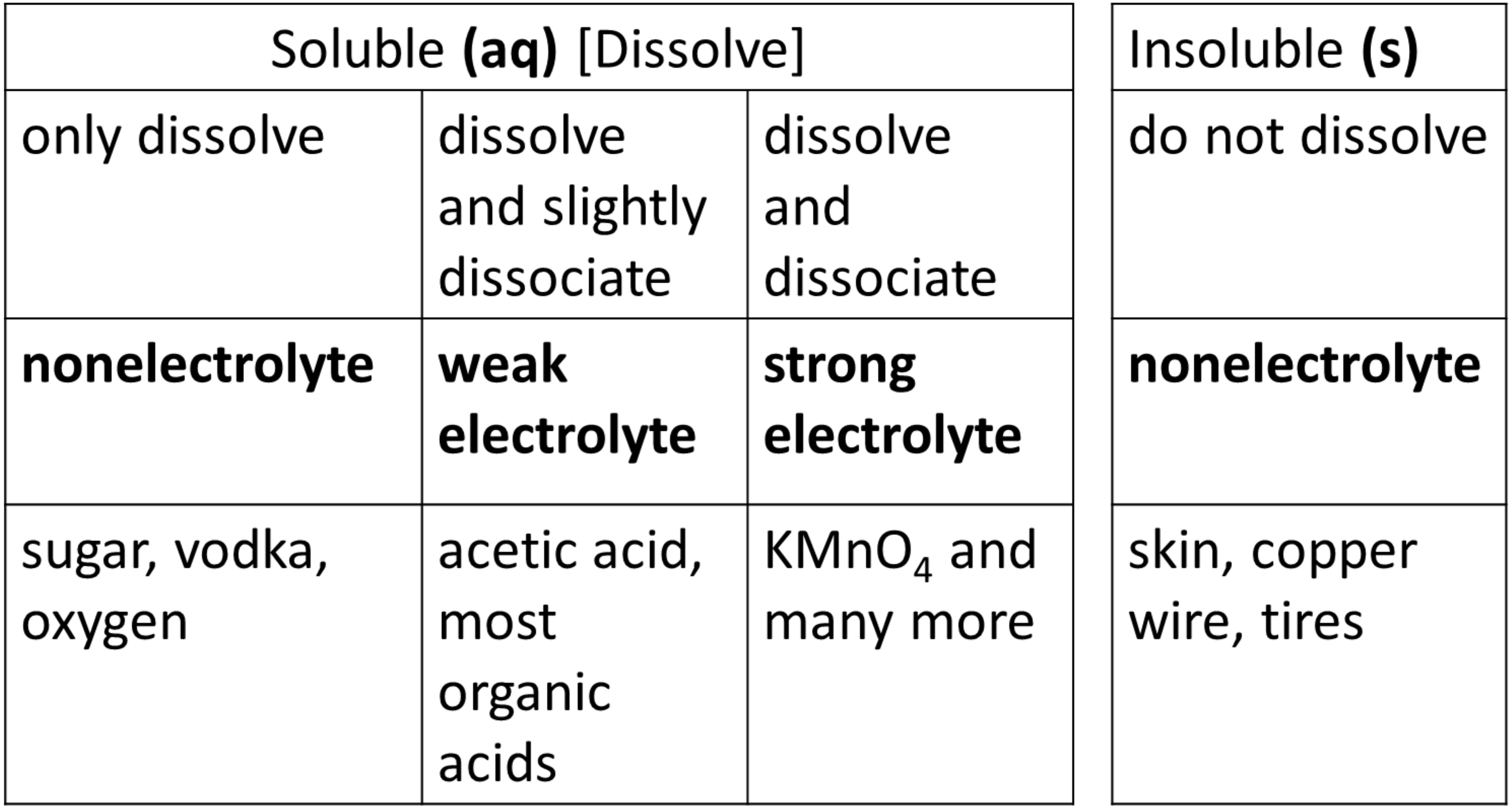
This net ionic equation is the net change taking place in the reaction when Ag^+ ions combine with Cl^- ions to form the precipitate solid silver chloride, AgCl.

Soluble vs. insoluble

- **Soluble** means all in the same phase, i.e., solid table salt dissolves in water.
  - Soluble compounds can be strong, weak, and nonelectrolytes; **these may or may not dissociate**.
  - **Soluble** substances **dissolve** in a solvent and form a homogeneous solution

- **Insoluble means** different phases, i.e., solid in a liquid.
  - Insoluble compounds are nonelectrolytes and do not dissociate.
  - **Insoluble** substances **do not dissolve** significantly in a solvent and form heterogeneous solutions

Soluble vs. insoluble



Soluble **(aq)** [Dissolve]			Insoluble **(s)**
only dissolve	dissolve and slightly dissociate	dissolve and dissociate	do not dissolve
nonelectrolyte	**weak electrolyte**	**strong electrolyte**	**nonelectrolyte**
sugar, vodka, oxygen	acetic acid, most organic acids	$KMnO_4$ and many more	skin, copper wire, tires

Soluble vs. insoluble

For info only

- **Soluble Compounds**

Most compounds of Group 1 elements and ammonium (NH_4^+) compounds

Chlorides (Cl^-), bromides (Cl^-), and iodides (I^-), *except* for those of Ag^+, Hg_2^{2+}, and Pb^{2+}

Nitrates (NO_3^-), acetates ($CH_3CO_2^-$), chlorates (ClO_3^-), and perchlorates (ClO_4^-)

Sulfates (SO_4^{2-}), *except* those of Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, Hg_2^{2+}, and Ag^+

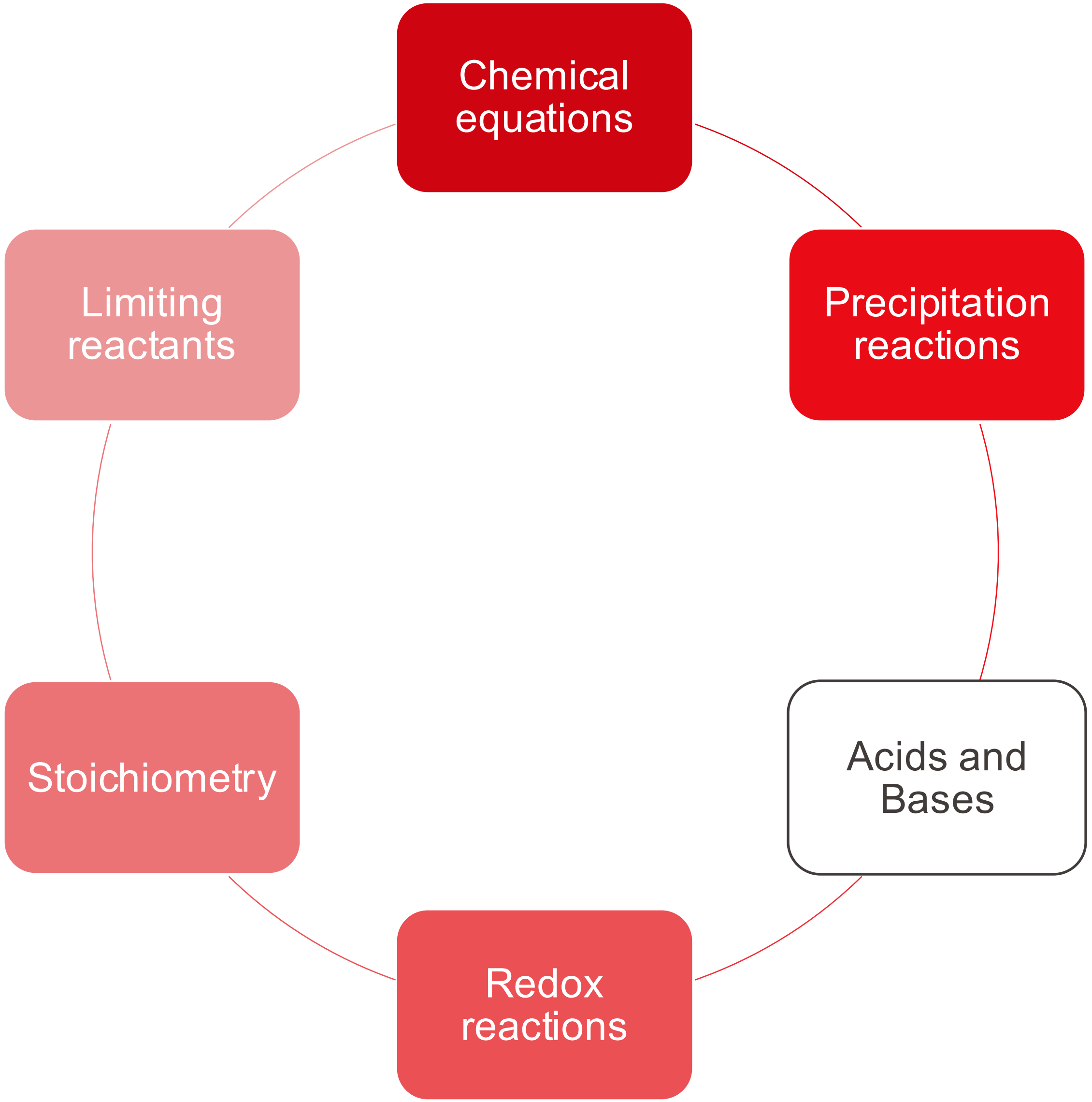
- **Insoluble Compounds**

Carbonates (CO_2^{2-}), chromates (CrO_4^{2-}), oxalates ($C_2O_4^{2-}$), and phosphates (PO_4^{3-}), *except* those of group 1 elements and NH_4^+

Sulfides (S^{2-}), *except* for those group 1and 2 elements and NH_4^+

Hydroxides (OH^-) and oxides (O^{2-}), *except* those of group 1 elements and 2 elements later than period 2.

Plan



Chemical equations

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Redox reactions

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Acids and bases

- **Acids** and **bases** change the color of certain dyes knowns as indicators

Andrew Lambert Photography/Science Source.

Acids and bases – definitions (Arrhenius, 1884)

- An **acid** is a compound that contains hydrogen and reacts with water to form hydrogen (H^+) ions.
  - Hydrochloric acid, HCl, is an **acid** because it produces hydrogen ion, H^+.

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

- A **base** is a compound that produces hydroxide ions (OH^-) in water.
  - Ammonia, NH_3, is a **base** because it produces hydroxide OH^-.

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

Acids and bases – definitions (Arrhenius, 1884)

- This definition has **limitations**:
  - The Arrhenius definition is *specific* to one solvent, *water*.
  - *Not all base reactions* produce the hydroxide ion, OH^-.
  - The key process in an acid and base reaction is a proton (H^+) *transfer* (little to do with OH^-).

Acids and bases – definitions (Brønsted-Lowry, 1923)

- Thomas Lowry in England and Johannes Brønsted in Denmark both developed the proton (H^+) transfer concept.
  - **Acid** – proton donor; **Base** – proton acceptor

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

  - HCl is a proton donor (**acid**); H_2O is a proton acceptor (**base**)
  - HCl transfers a hydrogen ion, H^+, to water, producing *hydronium ions* (H_3O^+) and chloride ions.
  - HCl is a Brønsted acid and H_2O is a Brønsted base in this reaction.

Identifying acids from their formula

- Acidic hydrogen atoms are usually written first in the formula.

  HCl
  HNO_3

- Organic acids contain carboxyl group, —COOH, usually written at the end of the formula. They transfer proton to form carboxylate ion, — COO^-

  CH_3COOH, acetic acid

- **Monoprotic acids** have one acidic proton; **polyprotic acids** can donate more than one proton from each molecule.

Strong and weak acids and bases

- **Strong acid** – completely deprotonated
- **Weak acid** – partially deprotonated
- **Strong base** – completely protonated.
- **Weak base** – partially protonated

Examples

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

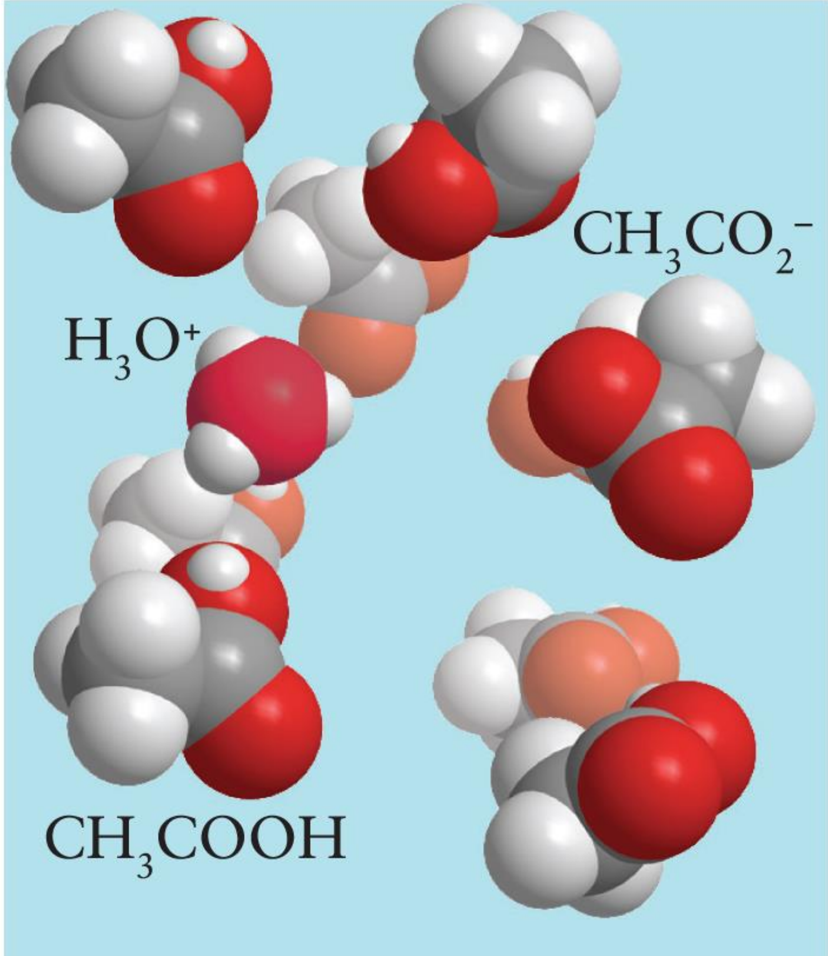
- Strong
- 100% proton transfer
- 100% deprotonation

$CH_3COOH\ (aq) + H_2O\ (l) \rightarrow CH_3COO^- \ (aq) + H_3O^+ \ (aq)$

Weak
~1% proton transfer
~1% deprotonation

Strong and weak acids and bases

- **Strong acids** deprotonate 100% and thus dissolve as ions in aqueous solution: HCl molecules are never found in solution.
- **Weak acids** deprotonate partially and thus dissolve primarily as molecules in solution. Acetic acid is found mainly as molecules in solution
- **Carboxylic acids** are weak acids



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Strongs and weak acids or bases

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Strong acids

- Hydrobromic acid, HBr (aq)
- Hydrochloric acid, HCl (aq)
- Hydroiodic acid, HI (aq)
- Nitric acid, HNO_3
- Perchloric acid, $HClO_4$
- Chloric acid, $HClO_3$
- Sulfuric acid, H_2SO_4 (to form HSO_4^-)

Strong bases

- Group 1 hydroxides
- Alkaline earth metal hydroxides ($Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$)
- Group 1 and Group 2 oxides

Neutralization reactions

- A reaction between an acid and a base is a **neutralization reaction**.

- Neutralization reactions occur between strong acids and metal hydroxides:

acid + metal hydroxide → salt + water

- “Salt” is taken from ordinary table salt, sodium chloride, but can refer to any ionic compound.

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

Neutralization reactions

$$\underset{\textbf{acid}}{2\,HNO_3\,(aq)} + \underset{\textbf{base}}{Ba(OH)_2\,(aq)} \rightarrow \underset{\textbf{a salt}}{Ba(NO_3)_2\,(aq)} + \underset{\textbf{water}}{2\,H_2O\,(l)}$$

Complete ionic equation:

$$2\,H^+\,(aq) + 2\,NO_3^-\,(aq) + Ba^{2+}\,(aq) + 2\,OH^-(aq) \rightarrow Ba^{2+}\,(aq) + 2\,NO_3^-\,(aq) + 2\,H_2O\,(l)$$

Remove spectators to obtain the net ionic equation:

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

Net reaction of any acid + strong base neutralization reaction is formation of water.

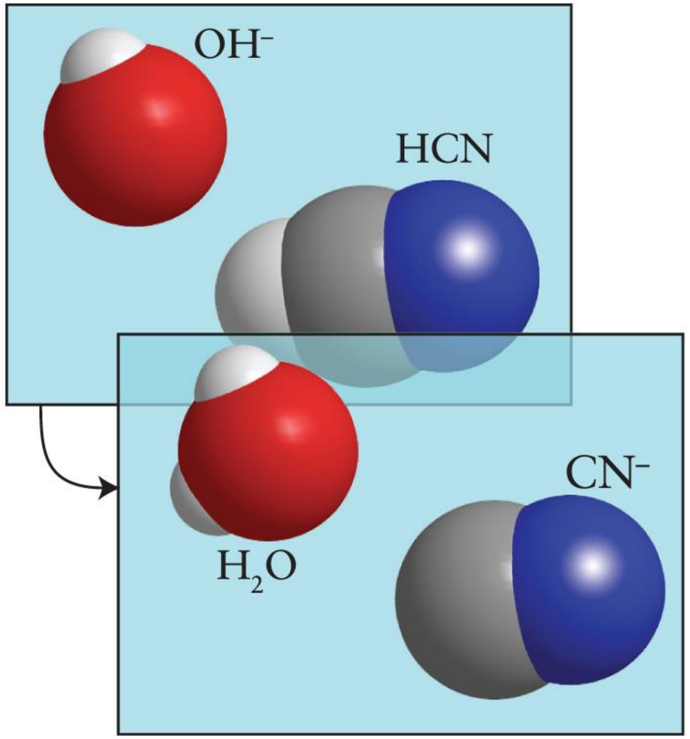
Neutralization reactions: weak acid neutralization

- HCN is a weak *acid*; only partially dissociates. Thus is HCN in solution and in net ionic equation.

- OH^- is proton acceptor – strong *base*

- Net ionic equation:

$$\mathbf{HCN}\ (aq) + \mathbf{OH^-}\ (aq) \rightarrow H_2O\ (l) + CN^-\ (aq)$$



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Neutralization reactions: weak base neutralization

- Weak base (partially protonates) + strong acid
- Molecular equation:

$$\mathbf{NH_3}\ (aq) + \mathbf{HCl}\ (aq) \rightarrow NH_4^+\ (aq) + Cl^-\ (aq)$$

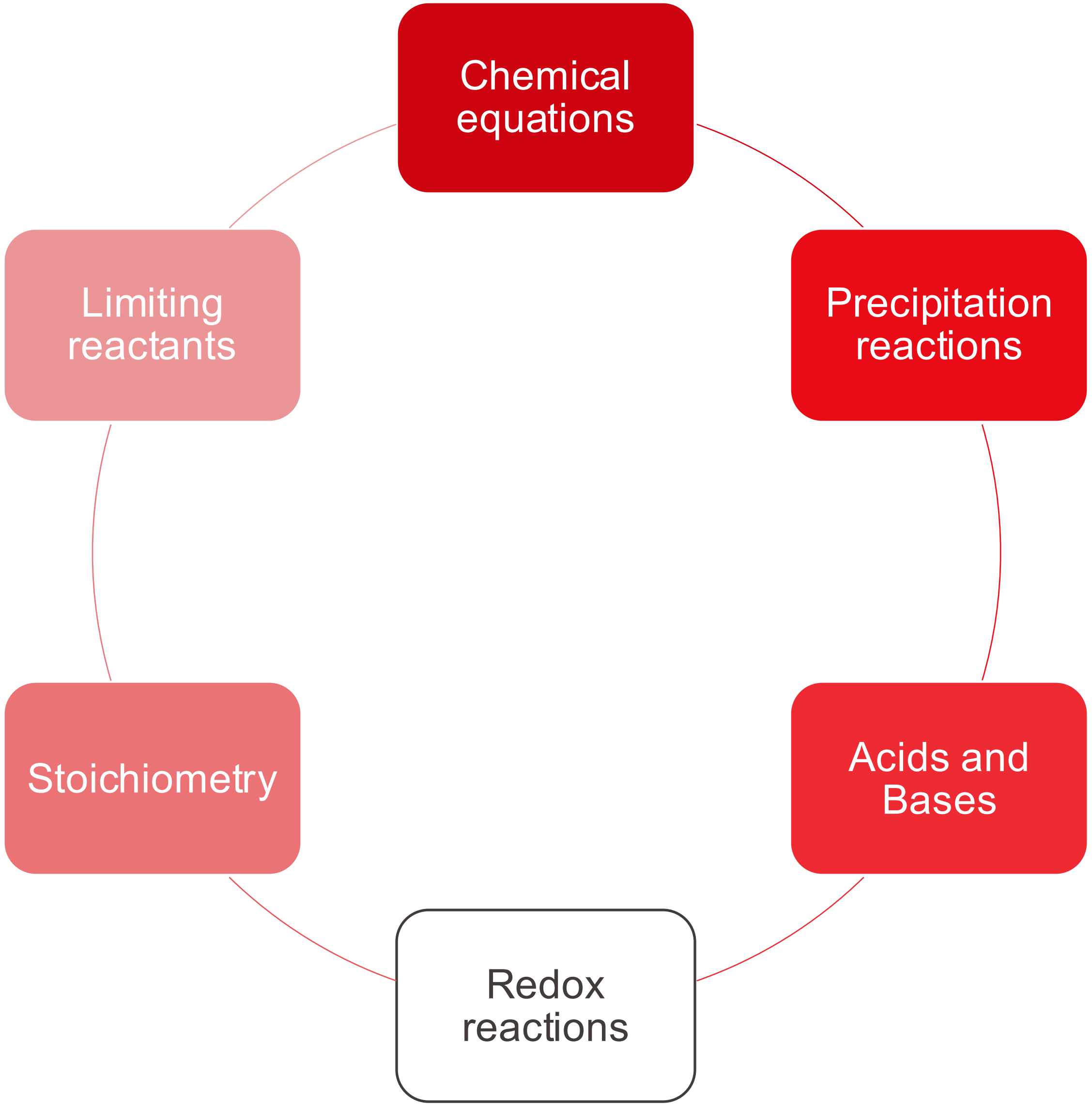
- Complete ionic equation:

$$NH_3\ (aq) + H^+\ (aq) + Cl^-\ (aq) \rightarrow NH_4^+\ (aq) + Cl^-\ (aq)$$

- Net ionic equation:

$$NH_3\ (aq) + H^+\ (aq) \rightarrow NH_4^+\ (aq)$$

Plan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Oxidation and reduction reactions

- Include:
  - Combustion reactions
  - Corrosion
  - Photosynthesis
  - Energy metabolism
  - Metal extraction

$$CO_2 + H_2O \xrightarrow{\text{sunlight}} \text{Glucose } (C_6H_{12}O_6) + O_2$$

Nigel Cattlin/Science Source.

Oxidation: loss of electrons

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- Classic meaning – reaction with oxygen, O_2

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

- Mg atoms lose electrons to form Mg^{2+} ions,
- O atoms in O_2 gain electrons to form O^{2-} ions.

$$2\ Mg + O_2 \rightarrow 2\ Mg^{2+} + 2\ O^{2-}$$

- **Oxidation** is the *loss* of electrons.
- Charge increases – oxidation

Oxidation: without oxygen

- When Mg reacts with chlorine gas, a similar reaction takes place.

$$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$$

- Mg atoms lose electrons to form Mg^{2+} ions (charge increases!)
- Cl atoms in Cl_2 gain electrons to form Cl^- ions (charge decreases!)

- Mg atoms *lose electrons* in an oxidation reaction even though no oxygen takes part.

$$Mg^0(s) + Cl_2(g) \rightarrow Mg^{2+}(s) + 2\ Cl^-(s)$$

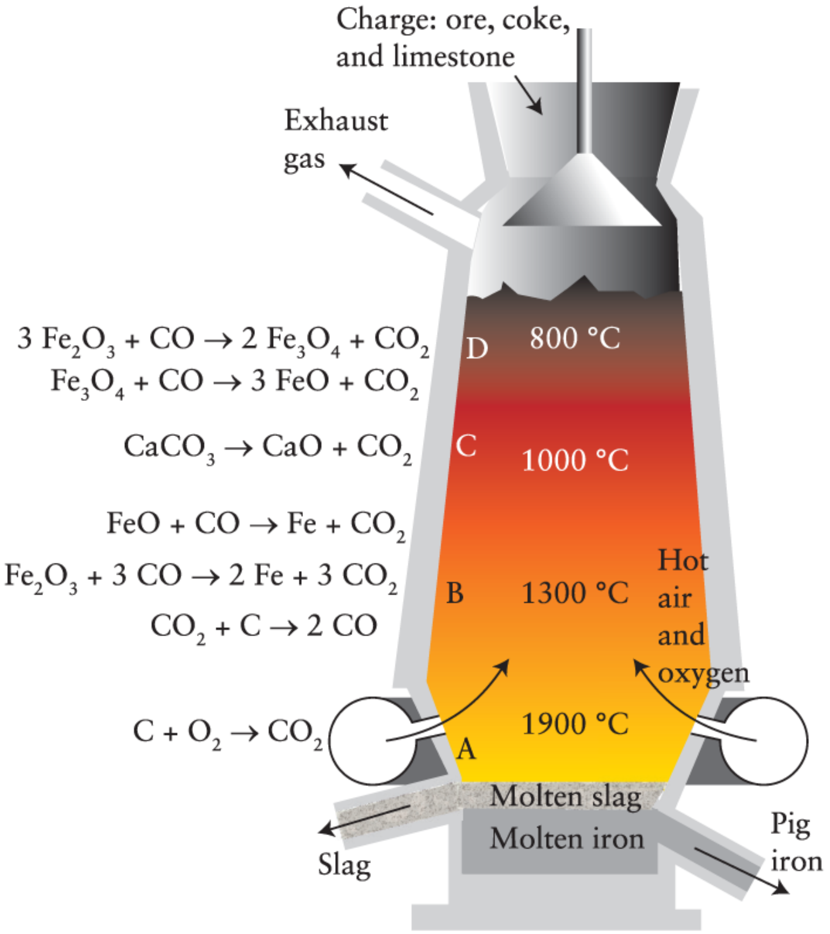
Reduction: gain of electrons

- **Reduction** is the *gain* of electrons

Example: reaction of iron(III) oxide and carbon monoxide in steel production

- Fe^{3+} ions *gain* electrons to form Fe^0 atoms
- Charge decreases – *reduction*

$$Fe_2O_3(s) + 3\,CO(g) \xrightarrow{\Delta} 2\,Fe(l) + 3\,CO_2(g)$$



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Oxidation and reduction occur simultaneously

$2\ NaBr(s) + Cl_2(s) \rightarrow 2\ NaCl(s) + Br_2(l)$

- Chlorine *is reduced* from 0 (in Cl_2) to −1 (in Cl^-). Chlorine *gained* electrons.
- Bromine *is oxidized* from −1 (in Br^-) to 0 (in Br_2). Bromine *lost* electrons.

➢ Electrons are transferred in redox (reduction-oxidation) reactions.

Oxidation numbers: keeping track of electrons

The **oxidation number** (also called **oxidation state**) is a number assigned to an atom in a chemical compound that represents the number of **electrons lost or gained** by that atom compared to its elemental state.

The oxidation number is defined so that:

- Oxidation is an increase in oxidation number.
- Reduction is a decrease in oxidation number.

$$2\ Cu^{+}(aq) + I_2(s) \rightarrow 2\ Cu^{2+}(aq) + 2\ I^{-}(aq)$$

- Copper's oxidation number increases from +1 to a +2.
- Iodine's oxidation number decreases from 0 to −1.

Rules for assigning an oxidation number (N_{ox})

- For uncombined elements, $N_{ox} = 0$.

  The oxidation numbers of the atoms in these elements are all zero: H_2, O_2, F_2, Cl_2, Li(s), U(s), C(s), P_4, S_8, etc.

- Sum of N_{ox} in ion or molecule = total charge.

  The N_{ox} of the ions in these compounds add to zero:

  $MgO(s) = Mg^{2+} + O^{2-}$

  $MgCl_2(s) = Mg^{2+} + 2\ Cl^{-}$

  $Fe_2O_3(s) = 2\ Fe^{3+} + 3\ O^{2-}$

  $NaBr(s) = Na^{+} + Br^{-}$

Oxidation numbers of elements in compounds

- Hydrogen
  - N_{ox} = +1 when H is combined with nonmetals
  - N_{ox} = −1 when H is combined with metals
- Groups 1 and 2
  - N_{ox} = group number

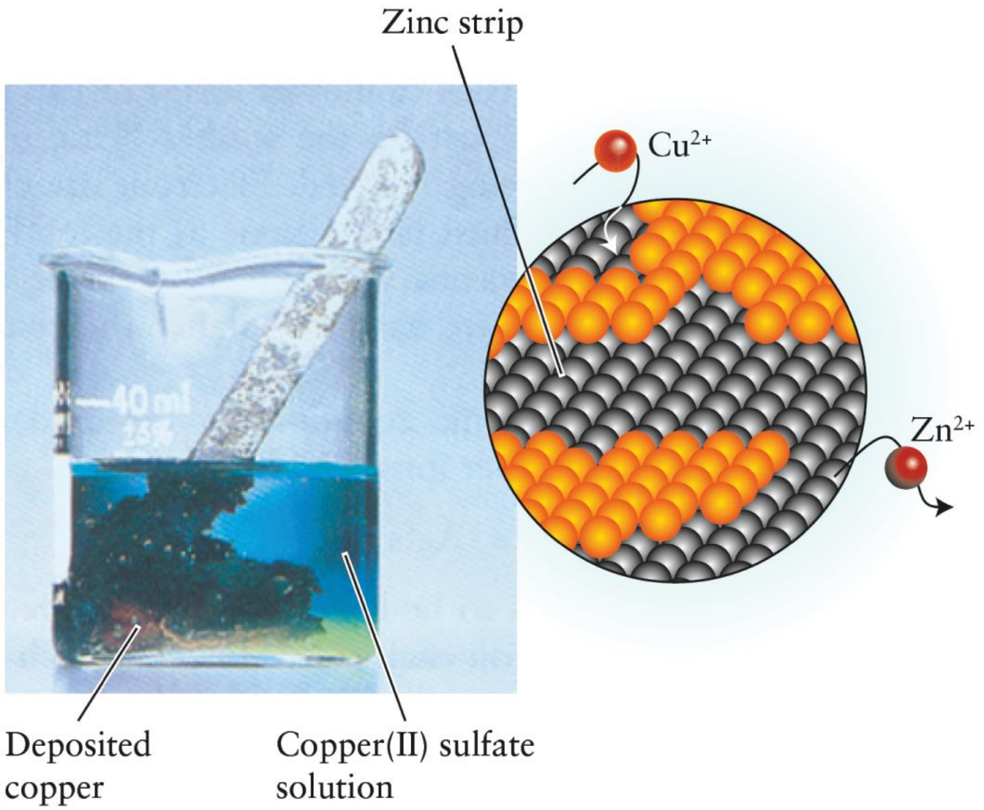
Oxidation numbers of elements in compounds

- Halogens
  - $N_{\mathrm{ox}} = -1$ unless combined with oxygen or a higher halogen.
  - Fluorine, $N_{\mathrm{ox}} = -1$ in all compounds
- Oxygen
  - $N_{\mathrm{ox}} = -2$ in most compounds
  - Exceptions: when combined with fluorine, peroxides (O_2^{2-}), superoxides (O_2^{-}), and ozonides (O_3^{-})

Oxidizing agents

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

- Zn loses electrons and is oxidized.
- Cu^{2+} gains the electrons that Zn loses.
- Cu^{2+} is the *agent* by which Zn is oxidized, Cu^{2+} is the **oxidizing agent**!



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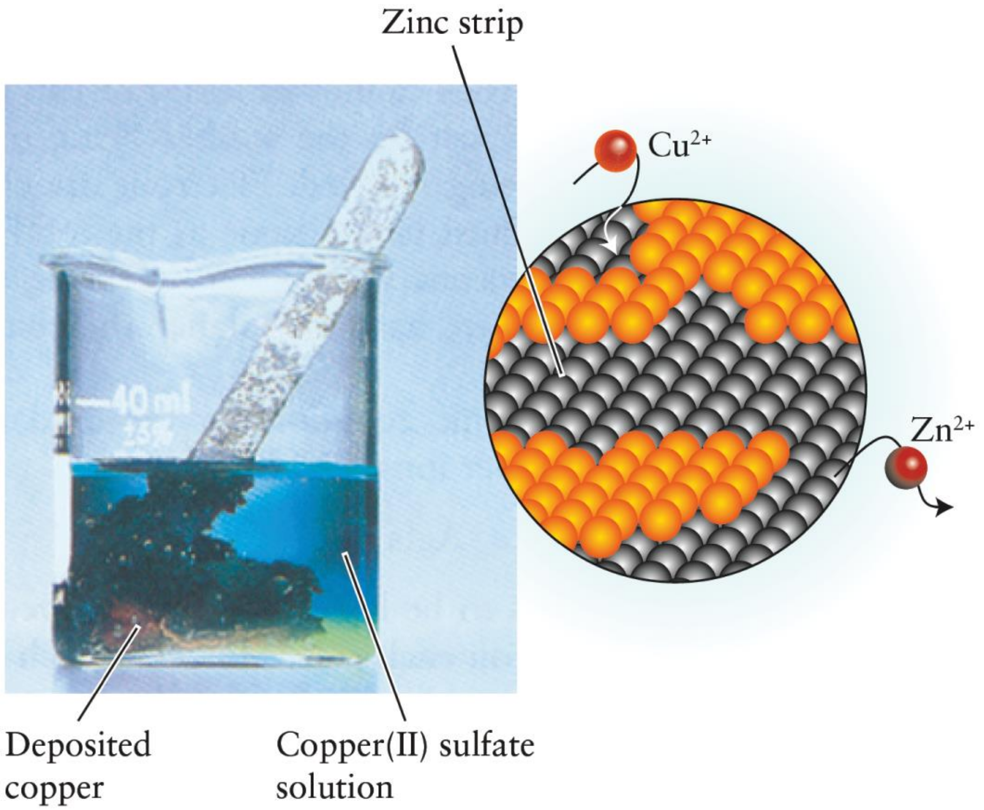
Reducing agents

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

- Zn loses electrons and is oxidized.
- Cu^{2+} gains the electrons that Zn loses.
- Zn is the *agent* by which Cu^{2+} is reduced, Zn is the **reducing agent**!



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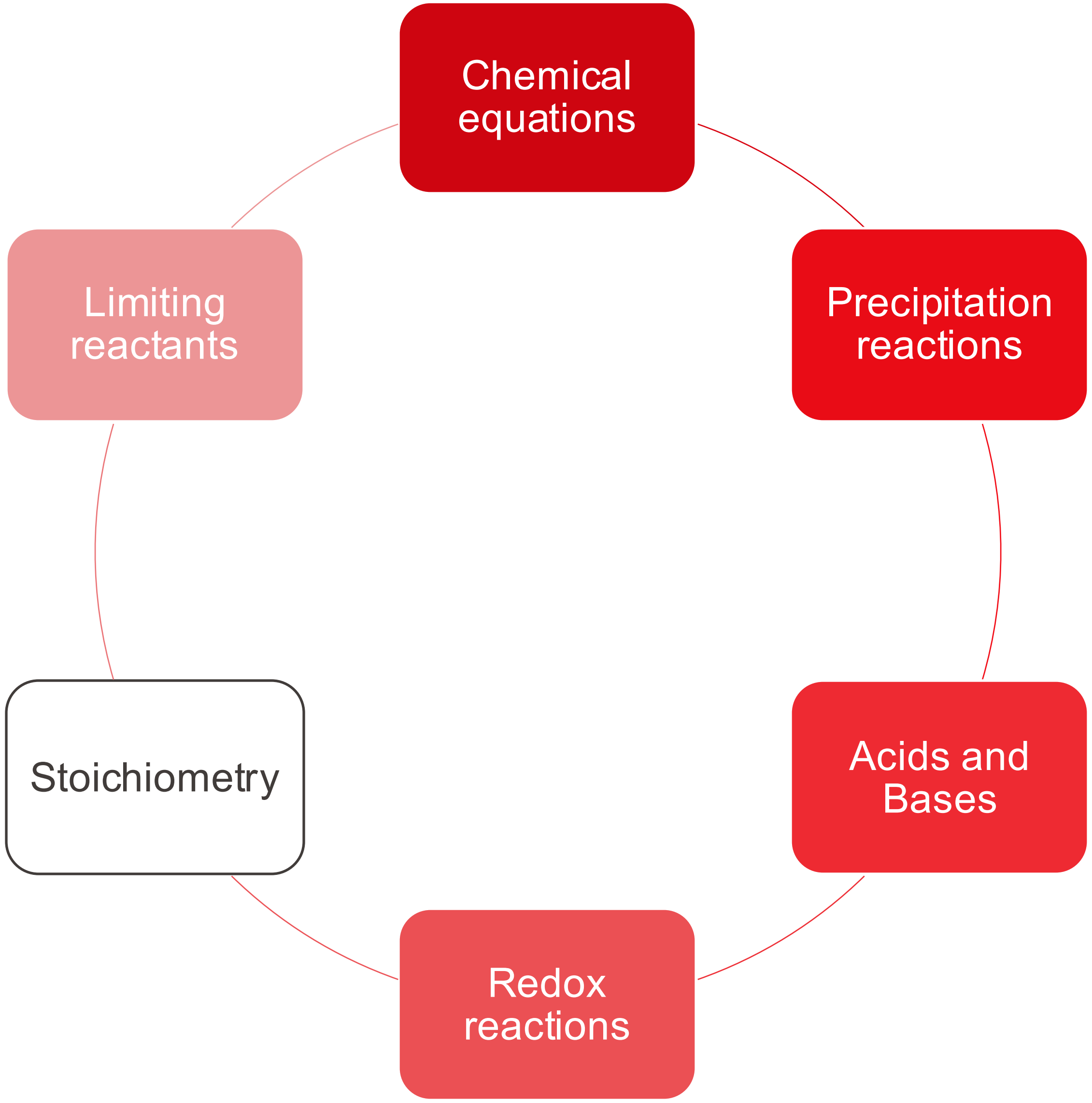
Plan



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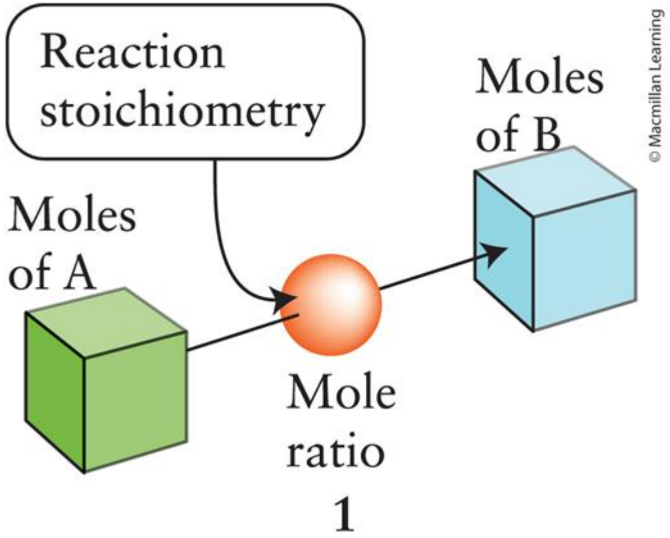
Reaction stoichiometry

- Answers questions:
  - *How much* product can we expect in a reaction?
  - *How much* reactant do we need?
- Is based on *balanced chemical equations*.
- Stoichiometric coefficients tell us the *relative number of moles* of reactants and products.
  - $\mathbf{1}\ N_2\ (g) + \mathbf{3}\ H_2\ (g) \rightarrow \mathbf{2}\ NH_3\ (g)$

Interpreting coefficients

$$\mathbf{1}\ N_2\ (g) + \mathbf{3}\ H_2\ (g) \rightarrow \mathbf{2}\ NH_3\ (g)$$

$$\mathbf{1}\ \text{mole}\ N_2\ (g) + \mathbf{3}\ \text{moles}\ H_2\ (g) \rightarrow \mathbf{2}\ \text{moles}\ NH_3\ (g)$$



- If 1 mol N_2 reacts, then 3 mol H_2 will be consumed and 2 mol NH_3 will be produced.

Practical importance

Predicting how much oxygen and hydrogen is needed on a space shuttle flight is a way to support life for the crew.

Argonne National Laboratory/Science Source.

Stoichiometry: mole-to-mole

How much water is formed when 0.25 mol O_2 reacts with hydrogen gas?

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$$

$$2\text{ mol } H_2(g) + 1\text{ mol } O_2(g) \rightarrow 2\text{ mol } H_2O(l)$$

The stoichiometric relation:

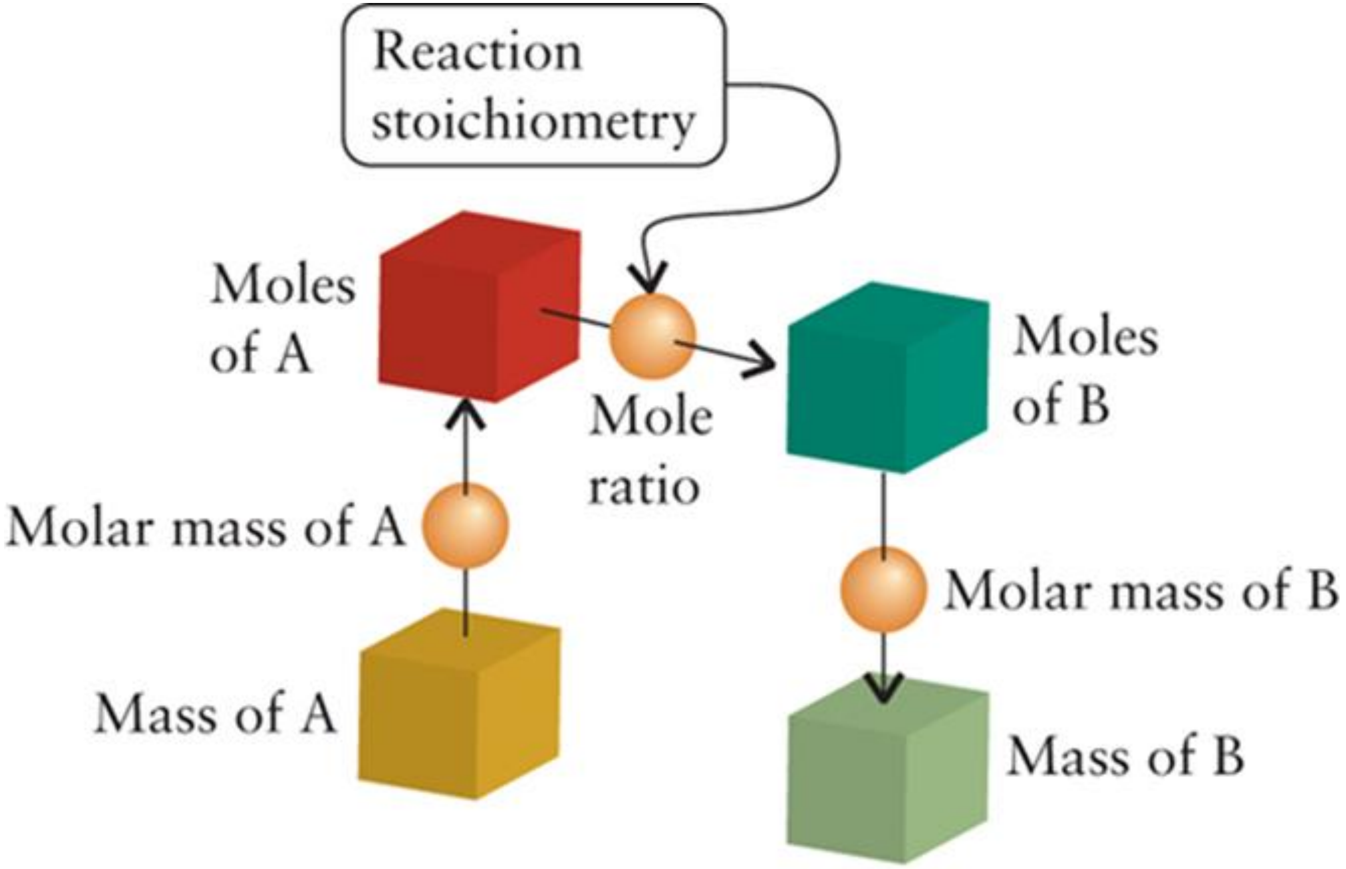
1 mol O_2 will form 2 mol H_2O

Use this relationship as a **mole ratio,** $\boxed{\dfrac{\text{Substance required}}{\text{Substance given}}}$

‘If I have’ $\dfrac{0.25\text{ mol } O_2}{1} \times \boxed{\dfrac{2\text{ mol } H_2O}{1\text{ mol } O_2}} = 0.50\text{ mol } H_2O$ ‘I will produce’

Stoichiometry: mass-to-mass

- Use molar mass to convert given mass to moles.
- Use a mole ratio to convert moles of given substance to moles of required substance.
- Use molar mass to convert moles to mass of required substance.

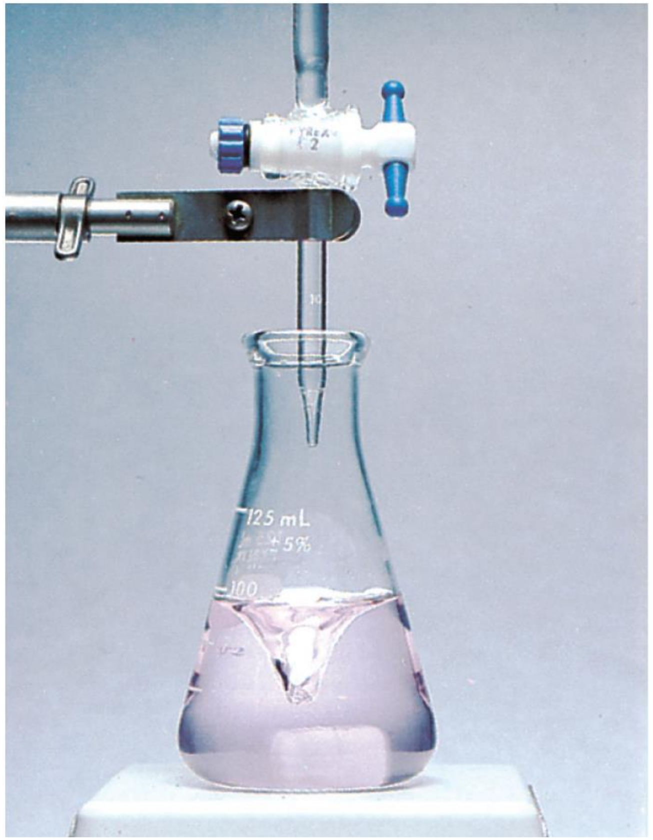


Volumetric analysis

Titration is a technique for finding concentrations.

- Titration involves two solutions, one of known concentration and one of unknown concentration.
- The unknown solution is called the **analyte**.
- A known volume is used.
- The known solution, or **titrant**, is added to the analyte solution until the stoichiometric point is reached.

An **indicator** is used to detect the stoichiometric point.



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Volumetric analysis - steps

1) Calculate moles of titrant used in titration:

$$n_{\text{titrant}} = c_{\text{titrant}} \times V_{\text{titrant}}$$

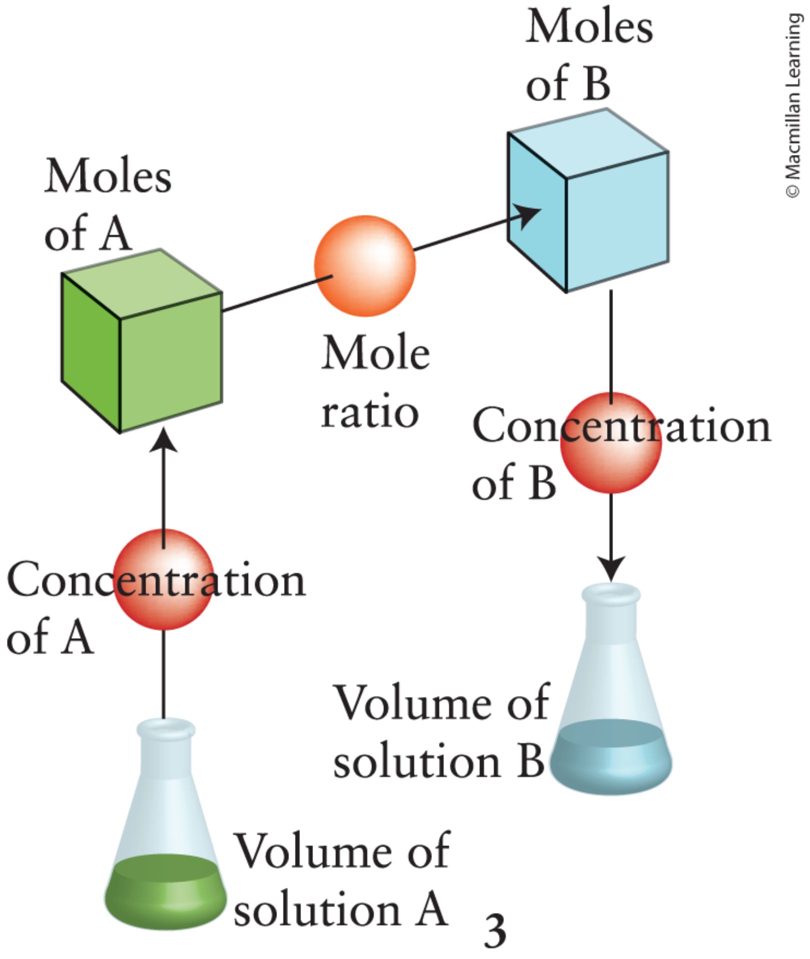
$$\left(c = \text{molar concentration} = \frac{\text{mol of solute}}{\text{liter of solution}}\right)$$

2) Calculate moles of analyte in titration:

$$n_{\text{analyte in titration}} = n_{\text{titrant}} \times \text{mole ratio}$$

3) Calculate molarity of analyte solution:

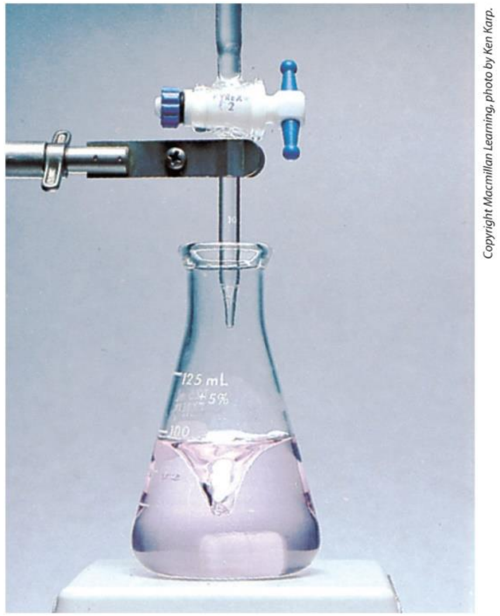
$$c_{\text{analyte solution}} = \frac{n_{\text{analyte in titration}}}{V_{\text{analyte in titration}}}$$



Determining the purity of a sample by redox titration

The amount of iron in an ore sample (=minerai) can be determined by titrating the ore with a potassium permanganate, $KMnO_4$, solution.

Ore is dissolved in hydrochloric acid, forming iron(II) ions, which react with MnO_4^-:



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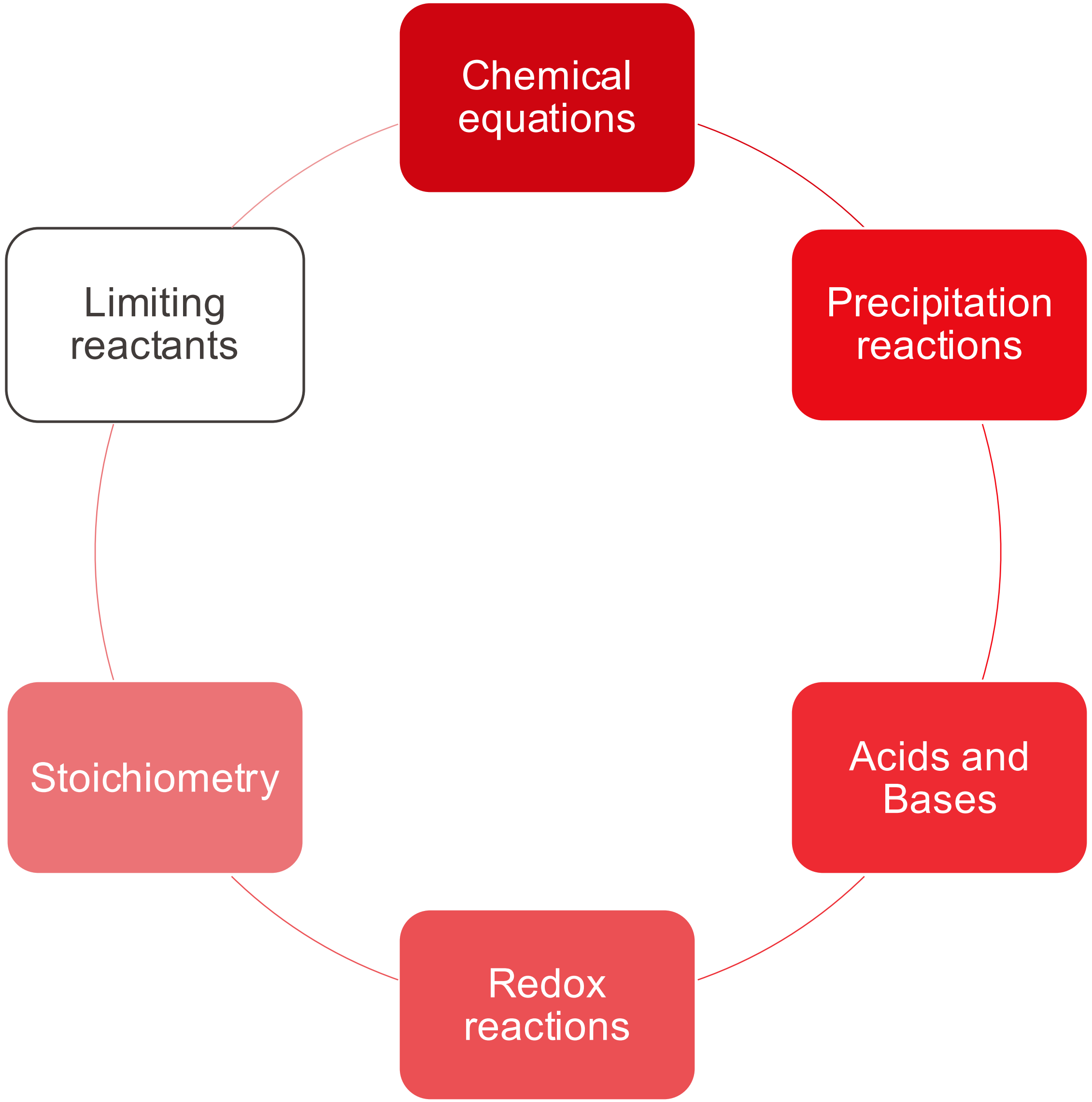
$$5\,Fe^{2+}(aq) + MnO_4^-(aq) + 8\,H^+(aq) \rightarrow 5\,Fe^{3+}(aq) + Mn^{2+}(aq) + 4\,H_2O(l)$$

The stoichiometric point is reached when all the Fe^{2+} has reacted and is detected as the purple color of the permanganate ion (see exercise session)

Plan


Chemical equations

Precipitation reactions

Acids and Bases

Redox reactions

Stoichiometry

Limiting reactants

Reaction yield

- **Actual yield** is the isolated quantity of a product that a chemist gathers after a reaction.
- A **theoretical yield** is the *maximum* quantity (amount, mass, or volume) of product possible in a reaction; it must be calculated.
- A **percentage yield** is the fraction of the theoretical yield actually produced, expressed as a percentage:

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

Limits of reactions

The **limiting reactant** governs the *maximum yield* of product.

Analogous to limited parts in a bicycle factory:

- 7 bike frames available
- 8 wheels available
- At 2 wheels per bike, only 4 bikes can be made.
- Wheels are the limiting “reactant”

Determining the limiting reactant

1. Determine the theoretical moles of product based on the amounts of each reactant provided.

2. The limiting reactant produces the *lowest* theoretical yield.

3. Use the theoretical moles of product from the limiting reactant in any further calculations.

Determining the limiting reactant

Suppose we start a reaction with 10 mol N_2 and 20 mol H_2; how much NH_3 can this make?

$$\mathbf{1}\ N_2\ (g) + \mathbf{3}\ H_2(g) \rightarrow \mathbf{2}\ NH_3(g)$$

Calculate the theoretical moles of NH_3 for each using the mole ratios.

$$\frac{10\text{ mol }N_2}{1} \times \frac{\mathbf{2}\text{ mol }NH_3}{\mathbf{1}\text{ mol }N_2} = 20\text{ mole }NH_3$$

$$\frac{20\text{ mol }H_2}{1} \times \frac{\mathbf{2}\text{ mol }NH_3}{\mathbf{3}\text{ mol }H_2} = 13\text{ mole }NH_3$$

Hydrogen produces *fewer* moles of product and is the *limiting* reactant.

Determining the limiting reactant

Method 1 – useful for two reactants

Use the mole ratios to determine if there is enough of one reactant to react with another.

Method 2 – useful for more than two reactants

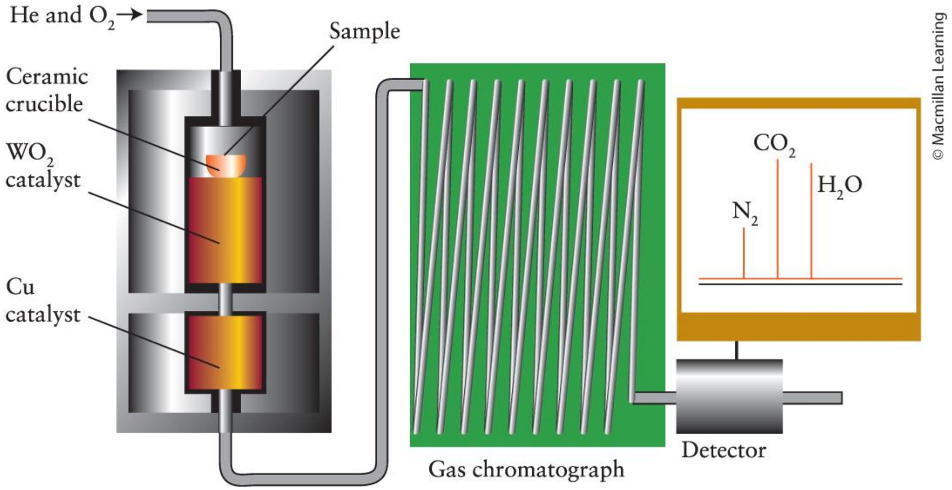
Calculate the theoretical yield of product (in moles) from each reactant separately to find which is smallest.

Combustion analysis for empirical formula

$$C_wH_xN_yO_z + O_2(g) \rightarrow CO_2(g) + H_2O(l) + N_2(g)$$

For information only

- O_2/He carrier gas ensures that sample = limiting reactant
- WO_2 catalyst converts any CO to CO_2.
- Cu removes excess O_2
- Gas chromatograph separates N_2, CO_2, and H_2O.
- Detector determines mass of N_2, CO_2, and H_2O.



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Combustion analysis for empirical formula

See exercices

$C_wH_xN_yO_z + O_2(g) \rightarrow CO_2(g) + H_2O(l) + N_2(g)$

- Use the stoichiometric relations to find the masses of carbon, nitrogen and hydrogen atoms in the sample.
  - 1 mol of CO_2 is chemically equivalent to 1 mol of C
  - 1 mol of H_2O is chemically equivalent to 2 mol of H
  - 1 mol of N_2 is chemically equivalent to 2 mol N

- Subtract these masses from the sample mass to obtain mass of O in the sample and convert to moles of O.
- Determine empirical formula from moles of C, H and O.

Have a beautiful day !

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