Roadmap

- Calculate amounts of reactants and products using stoichiometry and molarity
- Titrate to equivalence point
- Recognize, classify, write, balance, and predict product-favored redox reactions

Syllabus Learning Outcomes : 8, 9, and 10

SOLUTION STOICHIOMETRY

- Zinc reacts with acids to produce H_2 gas.
- Have 10.0 g of Zn
- What volume of 2.50M HCl is needed to convert all the Zn to H_2?

PLAN FOR STOICHIOMETRY CALCULATIONS

Mass zinc → Moles zinc → (Stoichiometric factor) → Moles HCl → Volume HCl

Moles HCl → Mass HCl

What volume of 2.50 M HCl converts 10.0g of Zn to H_2 gas?

What volume of 2.50 M HCl converts 10.0g of Zn to H_2 gas?

Step 1: Write the balanced equation

$Zn(s) + 2\,HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Step 2: Calculate amount of Zn

$$10.0 \text{ g Zn} \cdot \frac{1.00 \text{ mol Zn}}{65.39 \text{ g Zn}} = 0.153 \text{ mol Zn}$$

What volume of 2.50 M HCl converts 10.0g of Zn to H_2 gas?

Step 3: Use the stoichiometric factor

$$0.153 \text{ mol Zn} \cdot \frac{2 \text{ mol HCl}}{1 \text{ mol Zn}} = 0.306 \text{ mol HCl}$$

Step 4: Calculate volume of HCl

$$0.306 \text{ mol HCl} \cdot \frac{1.00 \text{ L}}{2.50 \text{ mol}} = 0.122 \text{ L HCl}$$

Titration

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- Combine two reactants to reach a *stoichiometric proportion* or endpoint guided by an indicator.
- Analyze the *analyte* (moles, grams, percentage, or concentration).
- Works for any reaction type

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Titrate an acid with a base

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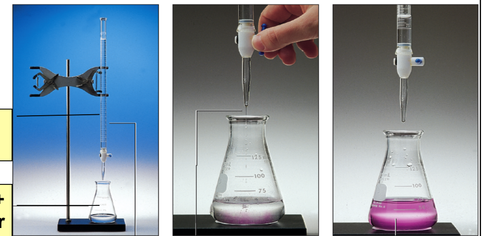
$H_2C_2O_4(aq) + 2NaOH(aq) \rightarrow Na_2C_2O_4(aq) + 2H_2O(\ell)$

titrant in buret

sample + indicator in flask



(a) Buret containing aqueous NaOH of accurately known concentration.

(b) A solution of NaOH is added slowly to the sample being analyzed.

(c) When the amount of NaOH added from the buret exactly equals the amount of H^+ supplied by the acid being analyzed, the dye (indicator) changes color.

Setup for titrating an acid with a base

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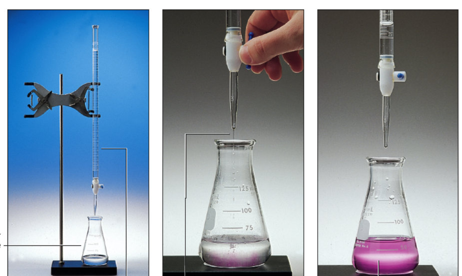
Flask containing aqueous solution of sample being analyzed



(a) Buret containing aqueous NaOH of accurately known concentration.

(b) A solution of NaOH is added slowly to the sample being analyzed.

(c) When the amount of NaOH added from the buret exactly equals the amount of H^+ supplied by the acid being analyzed, the dye (indicator) changes color.

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Titrations (cont' d)

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- In a titration, one reactant (the ***titrant***) is placed in a buret. The other reactant is placed in a flask along with a few drops of an indicator.
- The titrant is slowly added to the contents of the flask until the indicator changes color (the ***endpoint***).
- If the indicator has been chosen properly, the endpoint tells us when the reactants are present in stoichiometric proportion.
- A titration may be based on any of the previously discussed types of reactions …

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Titration

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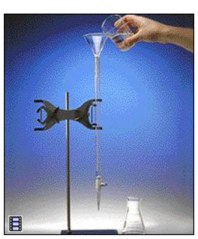

1. **Add solution from the buret.**
2. **Reagent (base) reacts with compound (acid) in solution in the flask.**
3. **Indicator shows when exact stoichiometric reaction has occurred.**
4. **Net ionic equation**

   $H^+ + OH^- \text{-->} H_2O$
5. **At equivalence point**

   moles H^+ = moles OH^-

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What is the concentration of NaOH given that 1.065 g of $H_2C_2O_4$ (oxalic acid) requires 35.62 mL of NaOH to titrate to an equivalence point?

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At equivalence point, moles H^+ = moles OH^-

Indicator solution changes color

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What is the concentration of the NaOH given that 1.065 g of $H_2C_2O_4$ (oxalic acid) requires 35.62 mL of NaOH for titration to an equivalence point?

Step 1: Calculate amount of $H_2C_2O_4$

$$1.065\ \text{g} \cdot \frac{1\ \text{mol}}{90.04\ \text{g}} = 0.0118\ \text{mol}$$

Step 2: Calculate amount of NaOH needed

$$0.0118\ \text{mol acid} \cdot \frac{2\ \text{mol NaOH}}{1\ \text{mol acid}} = 0.0236\ \text{mol NaOH}$$

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What is the concentration of the NaOH given that 1.065 g of $H_2C_2O_4$ (oxalic acid) requires 35.62 mL of NaOH for titration to an equivalence point?

Step 3: Calculate concentration of NaOH

$$\frac{0.0236\ \text{mol NaOH}}{0.03562\ \text{L}} = 0.663\ \text{M}$$

M_{NaOH} = [NaOH] = 0.663 M

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76.80g of apple requires 34.56mL of 0.663M NaOH for titration. What is mass % of malic acid?

Apples contain malic acid, $C_4H_6O_5$.

Titrate apples using standardized NaOH.

$C_4H_6O_5(aq) + 2\ NaOH(aq) \rightarrow$

$Na_2C_4H_4O_5(aq) + 2\ H_2O(\ell)$

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76.80g of apple requires 34.56mL of 0.663M NaOH for titration. What is mass % of malic acid?

Step 1: Calculate amount of NaOH used.

C • V = (0.663 M)(0.03456 L)

= 0.0229 mol NaOH

Step 2: Calculate amount of acid titrated.

$$0.0229\ \text{mol NaOH} \cdot \frac{1\ \text{mol acid}}{2\ \text{mol NaOH}}$$

= 0.0115 mol acid

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76.80g of apple requires 34.56mL of 0.663M NaOH for titration. What is mass % of malic acid?

Step 3: Calculate mass of acid titrated.

$$0.0115\ \text{mol acid} \cdot \frac{134\ \text{g}}{\text{mol}} = 1.54\ \text{g}$$

Step 4: Calculate % malic acid.

$$\frac{1.54\ \text{g}}{76.80\ \text{g}} \cdot 100\% = 2.01\%$$

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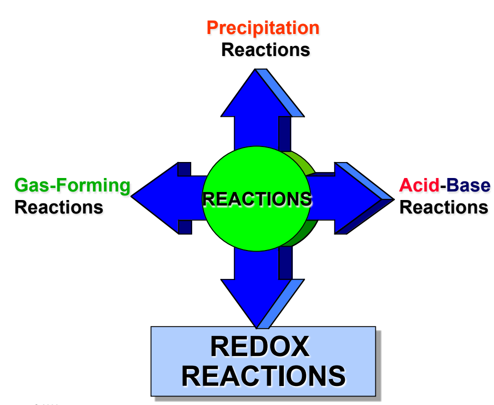

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Redox is Important ($)

Batteries

Corrosion

Manufacturing metals

Fuels

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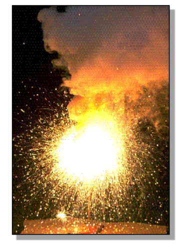
Oxidation-Reduction Reactions

Thermite reaction

$Fe_2O_3(s)$ + 2 Al(s)

---->

2 Fe(s) + $Al_2O_3(s)$



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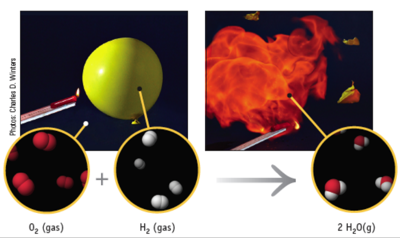
REDOX REACTIONS

REDOX = reduction & oxidation

2 $H_2(g)$ + $O_2(g)$ ---> 2 $H_2O(liq)$



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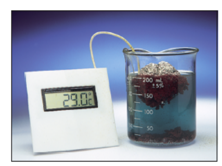
REDOX REACTIONS

REDOX = reduction & oxidation

Corrosion of aluminum

(a)



(b)

2 Al(s) + 3 $Cu^{2+}(aq)$ ---> 2 $Al^{3+}(aq)$ + 3 Cu(s)

LEO goes GER

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Redox Reactions Have Electron Transfer

Transfer leads to—

1. Increase in oxidation number of some element = OXIDATION
2. Decrease in oxidation number of some element = REDUCTION

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A REDOX REACTION -

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

If a species was oxidized then a species was reduced

LEO goes GER

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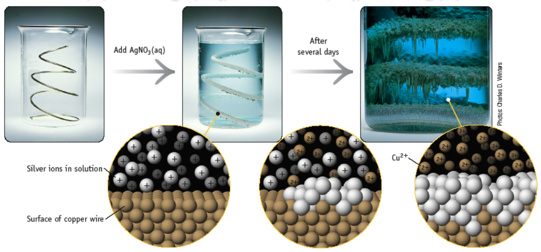
REDOX REACTIONS

$Cu(s) + 2\ Ag^{+}(aq) \text{ ---> } Cu^{2+}(aq) + 2\ Ag(s)$



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Determine Oxidation Numbers

The electric charge an element APPEARS to have when electrons are counted by some arbitrary rules:

1. Each atom in free element has oxidation # = 0
   Zn O_2 I_2 S_8 C_{60}

2. In simple ions, oxidation # = charge on ion.
   -1 for Cl^- +2 for Mg^{2+}

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Determine Oxidation Numbers

3. In compounds, oxidation # of O = -2
   (except in peroxides: in H_2O_2, O = -1)

4. In compounds, oxidation # of H = +1
   (except when H is associated with a metal as in NaH where it is -1)

χ

5. Algebraic sum of oxidation numbers
   = 0 for a compound
   = overall charge for an ion

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Determine Oxidation Numbers

NH_3	N =
ClO^-	Cl =
H_3PO_4	P =
MnO_4^-	Mn =
$NiCr_2O_7$	Cr =
C_3H_8	C =

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Balance a Redox Reaction

Corrosion of aluminum

$2\ Al(s) + 3\ Cu^{2+}(aq) \text{ --> } 2\ Al^{3+}(aq) + 3\ Cu(s)$

$Al(s) \text{ --> } Al^{3+}(aq) + 3\ e^-$

- Oxidation # of Al increases as e^- are lost by the metal.
- Therefore, Al is OXIDIZED
- Al is the REDUCING AGENT in this balanced half-reaction.

LEO goes GER

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Balance a Redox Reaction

Corrosion of aluminum

$2\ Al(s) + 3\ Cu^{2+}(aq) \text{ --> } 2\ Al^{3+}(aq) + 3\ Cu(s)$

$Cu^{2+}(aq) + 2\ e^- \text{ --> } Cu(s)$

- Oxidation # of Cu decreases as e^- are gained.
- Therefore, Cu is REDUCED
- Cu is the OXIDIZING AGENT in this balanced half-reaction.

LEO goes GER

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Balance a Redox Reaction

Multiply half-reactions by coefficients to cancel the electrons.

$2\ Al(s) \longrightarrow 2\ Al^{3+}(aq) + 6\ e^-$

$3\ Cu^{2+}(aq) + 6\ e^- \longrightarrow 3\ Cu(s)$

$2\ Al(s) + 3\ Cu^{2+}(aq) \longrightarrow 2\ Al^{3+}(aq) + 3\ Cu(s)$

Final eqn. is balanced for mass and charge.

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Examples of Redox Reactions

Metal + halogen

$2\ Al + 3\ Br_2 \longrightarrow Al_2Br_6$

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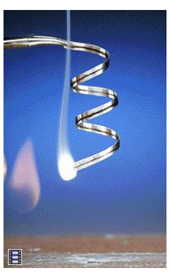
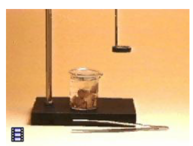
Examples of Redox Reactions





Nonmetal (P) + Oxygen $\longrightarrow P_4O_{10}$

Metal (Mg) + Oxygen $\longrightarrow MgO$

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Recognize Redox Reactions

Reaction Type	Oxidation	Reduction
In terms of oxygen	gain	loss
In terms of halogen	gain	loss
In terms of electrons	loss	gain
In terms of hydrogen	loss	gain

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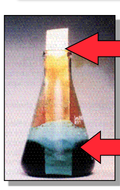
Common Oxidizing and Reducing Agents



Metals (Cu) are reducing agents

HNO_3 is an oxidizing agent

$Cu + 2NO_3^- + 4H^+ \rightarrow Cu^{2+} + 2NO_2 + 2H_2O$

Metals (Na, K, Mg, Fe) are reducing agents

$2K + 2H_2O \rightarrow 2KOH + H_2$

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Examples of Redox Reactions

Metal + acid

$Mg + HCl$

Mg = reducing agent

H^+ = oxidizing agent

Metal + acid

$Cu + HNO_3$

Cu = reducing agent

HNO_3 = oxidizing agent

Activity Series

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K
Ca
Mg
Al
Cr
Zn
Fe
Cd
Ni
Sn
Pb
H_2
Cu
Ag
Hg
Au

Metals above replace oxidized metals (in compounds) below

Strong reducing agent

Moderate

Weak reducing agent

Do Ni and Pb^{2+} react?
Do Cu and HCl react?

Balancing Equations for Redox Reactions

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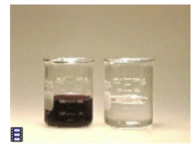
Some redox reactions have equations that must be balanced by special techniques.



MnO_4^- + 5 Fe^{2+} + 8 H^+ ---> Mn^{2+} + 5 Fe^{3+} + 4 H_2O

Mn = +7 Fe = +2 Mn = +2 Fe = +3

Balance Redox Reactions

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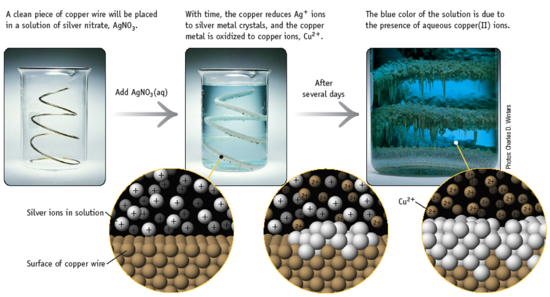
$Cu(s) + Ag^+(aq) \rightarrow Cu^{2+}(aq) + Ag(s)$



Balance Redox Reactions

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Step 1: Split the reaction into half-reactions, one for oxidation and the other for reduction.

Ox $Cu \rightarrow Cu^{2+}$

Red $Ag^+ \rightarrow Ag$

Step 2: Balance half reactions for mass (in aqueous acidic solution, can add H_2O to balance O and H^+ to balance H). **Already done in this case.**

Step 3: Balance half-reactions for charge by adding electrons.

Ox $Cu \rightarrow Cu^{2+} + 2e^-$

Red $Ag^+ + e^- \rightarrow Ag$

Balance Redox Reactions

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Step 4: Multiply half-reactions by factors so that the electrons cancel.

Reducing agent $Cu \rightarrow Cu^{2+} + 2e^-$

Oxidizing agent $2\ Ag^+ + 2\ e^- \rightarrow 2\ Ag$

Step 5: Add half-reactions (and simplify) to give the overall equation:

$Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$

The equation is balanced for both charge and mass.

Reduction of VO_2^+ with Zn

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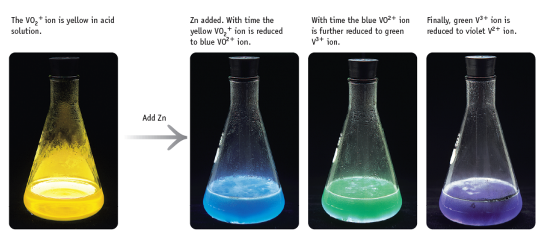

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Balance Redox Reactions

Balance the following in acid solution—

$VO_2^+ + Zn \rightarrow VO^{2+} + Zn^{2+}$

Step 1: Write the half-reactions

Ox $Zn \rightarrow Zn^{2+}$

Red $VO_2^+ \rightarrow VO^{2+}$

Step 2: Balance half-reactions for mass.

Ox $Zn \rightarrow Zn^{2+}$

Red $2H^+ + VO_2^+ \rightarrow VO^{2+} + H_2O$

Add H_2O on O-deficient side and add H^+ on other side for H-balance.

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Balance Redox Reactions

Step 3: Balance half-reactions for charge.

Ox $Zn \rightarrow Zn^{2+} + 2e^-$

Red $e^- + 2\,H^+ + VO_2^+ \rightarrow VO^{2+} + H_2O$

Step 4: Multiply by a factor to cancel e^-.

Ox $Zn \rightarrow Zn^{2+} + 2e^-$

Red $2e^- + 4H^+ + 2VO_2^+ \rightarrow 2VO^{2+} + 2H_2O$

Step 5: Add *balanced* half-reactions

$Zn + 4H^+ + 2VO_2^+ \rightarrow Zn^{2+} + 2VO^{2+} + 2H_2O$

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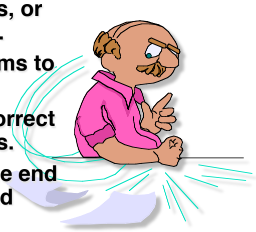
Tips to Balance Redox Rxns

- Never add O_2, O atoms, or O^{2-} to balance oxygen.
- Never add H_2 or H atoms to balance hydrogen.
- Be sure to write the correct charges on all the ions.
- Check your work at the end to make sure mass and charge are balanced.
- PRACTICE!



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Balance Redox Rxns in Base

- Write half reactions
- Balance half reactions for mass (H^+/H_2O)
- Balance half reactions for charge (e^-)
- Multiply by an appropriate factor to cancel out electrons
- Add half reactions
- Add OH^- to each side to neutralize all H^+ forming water
- Cancel out species that are the same on both sides

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End