CHAPTER 11

- Reactions in Aqueous Solutions II: Calculations

Chapter Goals

Aqueous Acid-Base Reactions

1. Calculations Involving Molarity
2. Titrations
3. The Mole Method and Molarity
4. Equivalent Weights and Normality

Oxidation-Reduction Reactions

5. The Half-Reaction Method
6. Adding in H^+, OH^- , or H_2O to Balance Oxygen or Hydrogen
7. Stoichiometry of Redox Reactions

Concentration of Solutions

- **Percent by mass**

$$\% \text{ by mass of solute} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

- **Molarity**

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

Calculations Involving Molarity

- Example: If 100.0 mL of 1.00 M NaOH and 100.0 mL of 0.500 M H_2SO_4 solutions are mixed, what will the concentration of the resulting solution be?
- What is the balanced reaction?
  - It is very important that we always use a balanced chemical reaction when doing stoichiometric calculations.

Calculations Involving Molarity

$$2NaOH \quad + \quad H_2SO_4 \rightarrow Na_2SO_4 \quad + \quad 2H_2O$$

strong base **strong acid**

	$2NaOH$	H_2SO_4	Na_2SO_4	$2H_2O$
Reaction Ratio:	2 mmol	1 mmol	1 mmol	2 mmol
Before Reaction:	100 mmol	50 mmol	0 mmol	0 mmol
After Reaction:	0 mmol	0 mmol	50 mmol	100 mmol

Calculations Involving Molarity

- What is the total volume of solution?

  100.0 mL + 100.0 mL = 200.0 mL

- What is the sodium sulfate amount, in mmol?

  50.0 mmol

- What is the molarity of the solution?

M = 50 mmol/200 mL = 0.250 M Na_2SO_4

Calculations Involving Molarity

- Example: If 130.0 mL of 1.00 M KOH and 100.0 mL of 0.500 M H_2SO_4 solutions are mixed, what will be the concentration of KOH and K_2SO_4 in the resulting solution?
- What is the balanced reaction?

Calculations Involving Molarity

	$2KOH$ +	H_2SO_4 $\rightarrow$	K_2SO_4 +	$2H_2O$
Reaction Ratio:	2 mmol	1 mmol	1 mmol	2 mmol
Before Reaction:	130 mmol	50 mmol	0 mmol	0 mmol
After Reaction:	30 mmol	0 mmol	50 mmol	100 mmol

Calculations Involving Molarity

- What is the total volume of solution?

  130.0 mL + 100.0 mL = 230.0 mL

- What are the potassium hydroxide and potassium sulfate amounts?

  30.0 mmol & 50.0 mmol

- What is the molarity of the solution?

M = 30.0 mmol/230.0 mL = 0.130 M KOH

M = 50.0 mmol/230.0 mL = 0.217 M K_2SO_4

Calculations Involving Molarity

- Example: What volume of 0.750 M NaOH solution would be required to completely neutralize 100 mL of 0.250 M H_3PO_4?

$$\mathbf{3NaOH + H_3PO_4 \rightarrow Na_3PO_4 + 3H_2O}$$

$$\text{?L NaOH} = 0.100 \text{ L } H_3PO_4 \times \frac{0.250 \text{ mol } H_3PO_4}{1 \text{ L } H_3PO_4} \times \frac{3 \text{ mol NaOH}}{1 \text{ mol } H_3PO_4} \times \frac{1 \text{ L NaOH}}{0.750 \text{ mol NaOH}} = 0.100 \text{ L NaOH}$$

Titrations

Acid-base Titration Terminology

1. **Titration** – A method of determining the concentration of one solution by reacting it with a solution of known concentration.
2. **Primary standard** – A chemical compound which can be used to accurately determine the concentration of another solution. Examples include KHP and sodium carbonate.
3. **Standard solution** – A solution whose concentration has been determined using a primary standard.
4. **Standardization** – The process in which the concentration of a solution is determined by accurately measuring the volume of the solution required to react with a known amount of a primary standard.

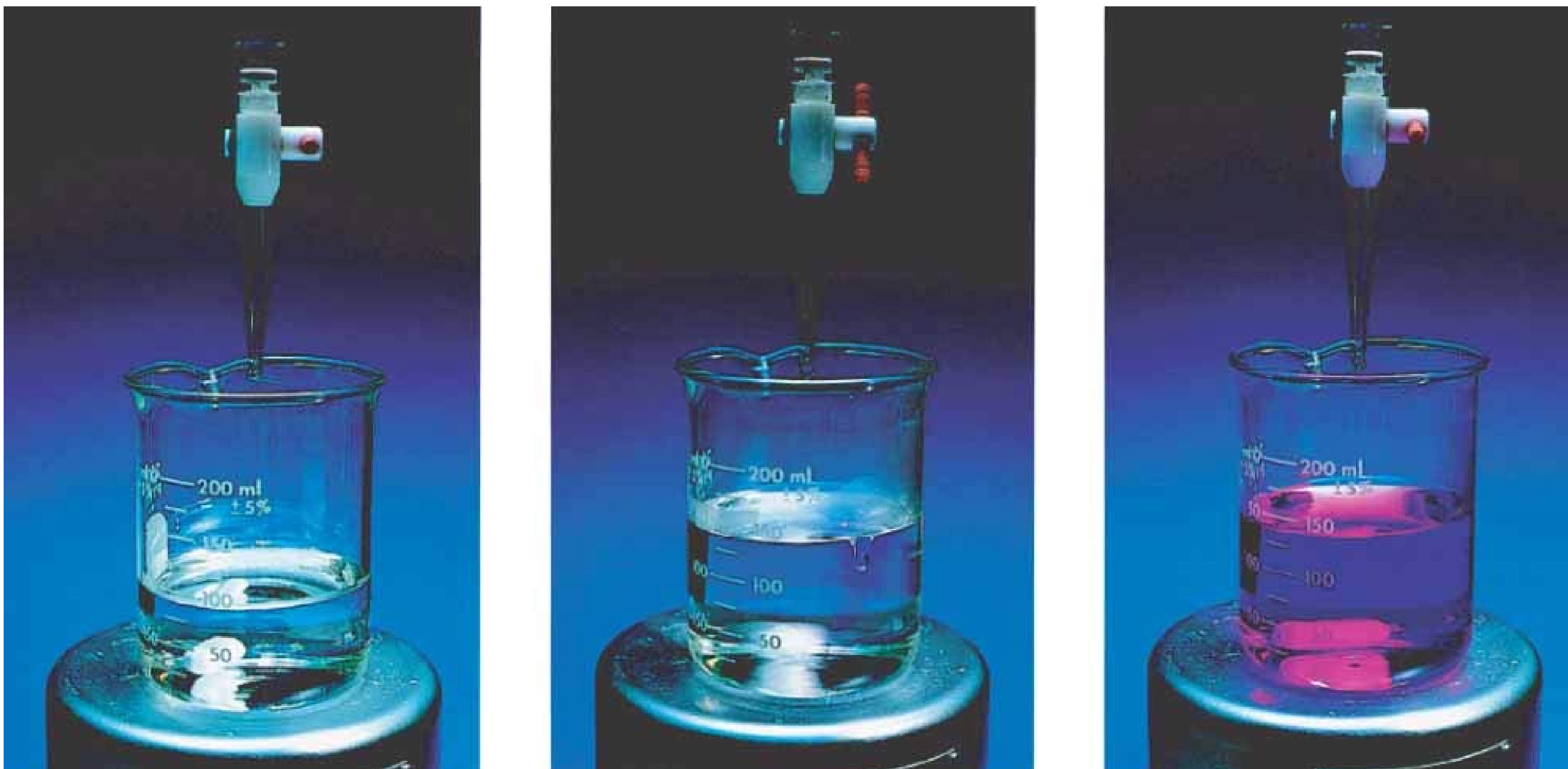
Titrations

Acid-base Titration Terminology

5. Indicator – A substance that exists in different forms with different colors depending on the concentration of the H^+ in solution. Examples are phenolphthalein and bromothymol blue.
6. Equivalence point – The point at which stoichiometrically equivalent amounts of the acid and base have reacted.
7. End point – The point at which the indicator changes color and the titration is stopped.

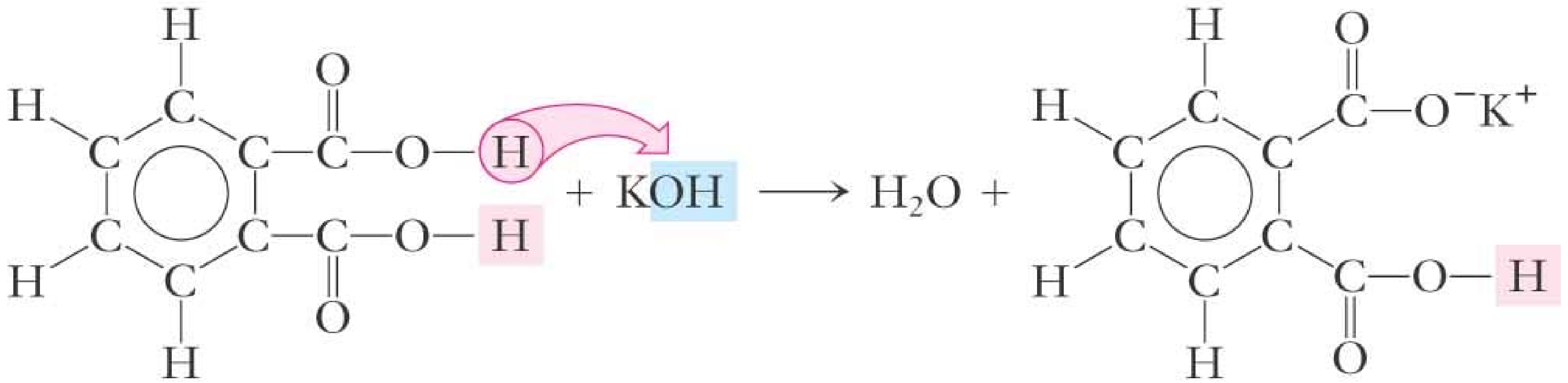
Titrations

- Acid-base Titration Terminology



The Mole Method and Molarity

- Potassium hydrogen phthalate is a very good primary standard. (Experiment 8 – Analysis of carbonated beverage)
  - It is often given the acronym, KHP.
  - KHP has a molar mass of 204.2 g/mol.



$C_6H_4(COOH)_2$
phthalic acid

$KC_6H_4(COO)(COOH)$
potassium hydrogen phthalate (KHP)

The Mole Method and Molarity

- Example: Calculate the molarity of a NaOH solution if 27.3 mL of it reacts with 0.4084 g of KHP.

$$\mathbf{NaOH \;+\; K\underline{H}P \rightarrow NaKP \;+\; H_2O}$$

$$\mathbf{?mol\ NaOH = 0.4084\ g\ KHP \times \frac{1\ mol\ KHP}{204.2\ g\ KHP} \times \frac{1\ mol\ NaOH}{1\ mol\ KHP} = 0.00200\ mol\ NaOH}$$

The Mole Method and Molarity

- Example: Calculate the molarity of a NaOH solution if 27.3 mL of it reacts with 0.4084 g of KHP.

$$\mathbf{NaOH + KHP \rightarrow NaKP + H_2O}$$

$$\mathbf{?M\ NaOH = \frac{0.00200\ mol\ NaOH}{0.0273\ L\ NaOH} = 0.0733\ M\ NaOH}$$

The Mole Method and Molarity

- **Example: Calculate the molarity of a sulfuric acid solution if 23.2 mL of it reacts with 0.212 g of Na_2CO_3.**

$$Na_2CO_3 \ + \ H_2SO_4 \rightarrow Na_2SO_4 \ + \ CO_2 \ + \ H_2O$$

$$?\text{mol } H_2SO_4 = 0.212 \text{ g } Na_2CO_3 \times \frac{1 \text{ mol } Na_2CO_3}{106 \text{ g } Na_2CO_3} \times$$

$$\frac{1 \text{ mol } H_2SO_4}{1 \text{ mol } Na_2CO_3} = 0.00200 \text{ mol } H_2SO_4$$

The Mole Method and Molarity

- **Example: Calculate the molarity of a sulfuric acid solution if 23.2 mL of it reacts with 0.212 g of Na_2CO_3.**

$$Na_2CO_3 + H_2SO_4 \rightarrow Na_2SO_4 + CO_2 + H_2O$$

$$?M\ H_2SO_4 = \frac{0.00200 \text{ mol } H_2SO_4}{0.0232 \text{ L } H_2SO_4} = 0.0862 \text{ M } H_2SO_4$$

Oxidation-Reduction Reactions

- We have previously gone over the basic concepts of oxidation & reduction in Chapter 4.
- Rules for assigning oxidation numbers were also introduced in Chapter 4.
  - Refresh your memory as necessary.
- We shall learn to balance redox reactions using the half-reaction method.

The Half-Reaction Method

Half reaction method rules:

1. **Write the unbalanced reaction.**
2. **Break the reaction into 2 half reactions:**

One oxidation half-reaction and

One reduction half-reaction

Each reaction must have complete formulas for molecules and ions.

3. **Mass balance each half reaction by adding appropriate stoichiometric coefficients. To balance H and O we can add:**
   - **H^+ or H_2O in acidic solutions.**
   - **OH^- or H_2O in basic solutions.**

The Half-Reaction Method

4. **Charge balance the half reactions by adding appropriate numbers of electrons.**
   - **Electrons will be products in the oxidation half-reaction.**
   - **Electrons will be reactants in the reduction half-reaction.**
5. **Multiply each half reaction by a number to make the number of electrons in the oxidation half-reaction equal to the number of electrons reduction half-reaction.**
6. **Add the two half reactions.**
7. **Eliminate any common terms and reduce coefficients to smallest whole numbers.**

The Half-Reaction Method

	To balance O:		To balance H:
In Acidic Solution:	Add H_2O	and then $\rightarrow$	Add H^+
In Basic Solution:	For *each* O needed, (1) add *two* OH^- to side needing O and (2) add *one* H_2O to other side	and then $\rightarrow$	For *each* H needed, (1) add *one* H_2O to side needing H and (2) add *one* OH^- to other side

The Half-Reaction Method

- **Example: Tin (II) ions are oxidized to tin (IV) by bromine. Use the half reaction method to write and balance the net ionic equation.**

Starting Reaction

$$Sn^{2+} + Br_2 \rightarrow Sn^{4+} + Br^-$$

Mass balance the half-reaction.

$$Sn^{2+} \rightarrow Sn^{4+}$$

Charge balance the half-reaction.

$$Sn^{2+} \rightarrow Sn^{4+} + 2e^-$$

Electrons are products thus this is the oxidation half-reaction

The Half-Reaction Method

Starting Reaction

$$Sn^{2+} + Br_2 \rightarrow Sn^{4+} + Br^-$$

$$Sn^{2+} \rightarrow Sn^{4+} + 2e^-$$

Mass balance the other half-reaction.

$$Br_2 \rightarrow 2\,Br^-$$

Charge balance the other half-reaction.

$$Br_2 + 2e^- \rightarrow 2\,Br^-$$

This is the reduction half reaction.

The Half-Reaction Method

$$Sn^{2+} + Br_2 \rightarrow Sn^{4+} + Br^- \quad \text{starting reaction}$$

Add the two half reactions.

$$Sn^{2+} \rightarrow Sn^{4+} + 2e^- \quad \text{ox. half reaction}$$

$$\underline{Br_2 + 2e^- \rightarrow 2\,Br^-} \quad \text{red. half reaction}$$

$$Sn^{2+} + Br_2 \rightarrow Sn^{4+} + 2\,Br^- \quad \text{balanced reaction}$$

The Half-Reaction Method

- **Example: Dichromate ions oxidize iron (II) ions to iron (III) ions and are reduced to chromium (III) ions in acidic solution. Write and balance the net ionic equation for the reaction.**

$$Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+} \text{ starting reaction}$$

The Half-Reaction Method

- **Example: Dichromate ions oxidize iron (II) ions to iron (III) ions and are reduced to chromium (III) ions in acidic solution. Write and balance the net ionic equation for the reaction.**

$$Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+} \quad \text{starting reaction}$$

Mass balance the half-reaction.

$$Fe^{2+} \rightarrow Fe^{3+}$$

Charge balance the half-reaction.

$$Fe^{2+} \rightarrow Fe^{3+} + 1e^{-}$$

Is this an oxidation or reduction half-reaction?

The Half-Reaction Method

$Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+}$ starting rxn

$Fe^{2+} \rightarrow Fe^{3+} + 1e^-$ oxidation

Mass balance the 2nd half - reaction.

$$Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+}$$

$Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+}$ starting rxn

$Fe^{2+} \rightarrow Fe^{3+} + 1e^-$ oxidation

Mass balance the 2nd half - reaction

$$Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + 7\,H_2O$$

The Half-Reaction Method

$$Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+} \qquad \text{starting rxn}$$

$$Fe^{2+} \rightarrow Fe^{3+} + 1e^{-} \qquad \text{oxidation}$$

$$\text{Mass balance the } 2^{nd} \text{ half - reaction}$$

$$14\,H^{+} + Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + 7\,H_2O$$

$$Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+} \qquad \text{starting rxn}$$

$$Fe^{2+} \rightarrow Fe^{3+} + 1e^{-} \qquad \text{oxidation}$$

$$\text{Mass balance the } 2^{nd} \text{ half - reaction}$$

$$14\,H^{+} + Cr_2O_7^{2-} \rightarrow 2\,Cr^{3+} + 7\,H_2O$$

$$\text{Charge balance the } 2^{nd} \text{ half - reaction.}$$

$$14\,H^{+} + Cr_2O_7^{2-} + 6\,e^{-} \rightarrow 2\,Cr^{3+} + 7\,H_2O$$

The Half-Reaction Method

$$Cr_2O_7^{2-} + Fe^{2+} \rightarrow Cr^{3+} + Fe^{3+}$$ start rxn

Add the two half-reactions.

$$6\left(Fe^{2+} \rightarrow Fe^{3+} + 1e^{-}\right)$$ ox.

$$\underline{Cr_2O_7^{2-} + 14\,H^{+} + 6\,e^{-} \rightarrow 2\,Cr^{3+} + 7\,H_2O}$$ red.

$$6\,Fe^{2+} + Cr_2O_7^{2-} + 14\,H^{+} \rightarrow 6\,Fe^{3+} + 2\,Cr^{3+} + 7\,H_2O$$

The Half-Reaction Method

- Example: In basic solution hydrogen peroxide oxidizes chromite ions, $Cr(OH)_4^-$, to chromate ions, CrO_4^{2-}. The hydrogen peroxide is reduced to hydroxide ions. Write and balance the net ionic equation for this reaction.

You do it!

$$\mathbf{Cr(OH)_4^- \quad + \quad H_2O_2 \quad \rightarrow \quad Cr_2O_4^{2-}}$$

The Half-Reaction Method

- Example: When chlorine is bubbled into basic solution, it forms hypochlorite ions and chloride ions. Write and balance the net ionic equation.

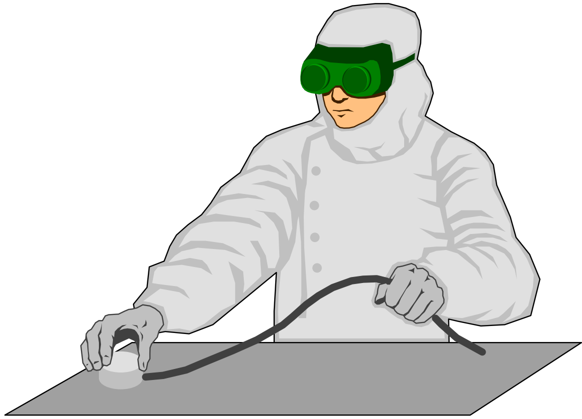
<u>You do it!</u>

- This is a disproportionation redox reaction. The same species, in this case Cl_2, is both reduced and oxidized.

$$\mathbf{Cl_2 \rightarrow ClO^- + Cl^- \quad (basic\ solution)}$$

End of Chapter 11

- Redox reactions are very important commercially.



Homework Assignment

One-line Web Learning (OWL):

Chapter 11 Exercises and Tutors –

Required by May 5th – 11:00 pm