

Reactions in Aqueous Solutions

The image shows a clear plastic IV drip chamber hanging from a metal stand on the left. To its right is a box of Opti-Flex Replenish sports drink, labeled 'VALUE PACK' and '20 SERVINGS'. The box is green and blue with a picture of the drink's container.

What are some of the driving forces that promote reactivity?

Reaction Type:	Driving Force
Precipitation	Formation of an insoluble compound
Acid-Base Neutralization	Formation of a salt and water (proton transfer)
Gas-Forming	Evolution of water-insoluble gas such as CO ₂
Oxidation-Reduction	Electron transfer

Solubility: What is the relative attraction of the particles in the compound for each other vs. water?

A water molecule is electrically positive on one side (the H atoms) and electrically negative on the other (the O atom). These charges enable water to interact with negative and positive ions in aqueous solution.

(a) Water surrounding a cation: Water molecules are oriented with their red oxygen atoms (labeled '-') towards the cation and their white hydrogen atoms (labeled '+') away from it.

(b) Water surrounding an anion: Water molecules are oriented with their white hydrogen atoms (labeled '+') towards the anion and their red oxygen atoms (labeled '-') away from it.

Copper chloride is added to water. Interactions between water and the Cu²⁺ and Cl⁻ ions allow the solid to dissolve. The ions are now sheathed in water molecules.

Hydrated Ion

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Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
<p>CuCl₂ Cu²⁺ Cl⁻</p>	<p>Acetic acid Acetate ion H⁺</p>	<p>Ethanol</p>
<p style="text-align: center;">Showing the hydration process</p>		

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The double arrow signifies the partial dissociation of the acetic acid

$$\text{CH}_3\text{CO}_2\text{H}(\text{aq}) \rightleftharpoons \text{CH}_3\text{CO}_2^-(\text{aq}) + \text{H}^+(\text{aq})$$

acetic acid
< 5% ionized
= weak electrolyte

acetate ion

hydrogen ion

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Compounds of hydroxide are mostly insoluble in water: (From Left to Right): NaOH, Ca(OH)₂, Fe(OH)₃, Ni(OH)₂

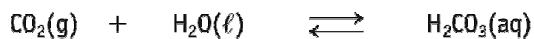
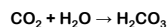
The image shows four test tubes. From left to right: 1. Clear colorless liquid (NaOH). 2. Slightly cloudy white liquid (Ca(OH)₂). 3. Dark brown precipitate at the bottom (Fe(OH)₃). 4. Green precipitate at the bottom (Ni(OH)₂).

Later you will be quantifying the solubility using solubility product values

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Oxides of nonmetals (acidic anhydrides) tend to produce acids in water:

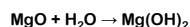
Ex.



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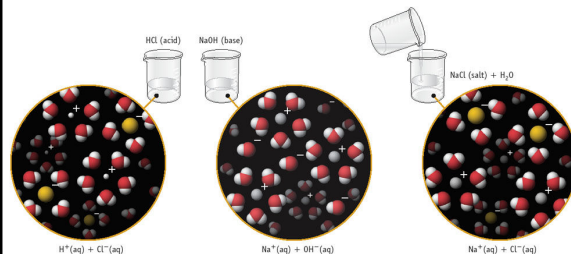
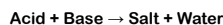
Oxides of metals (basic anhydrides) tend to produce bases in water:

Ex.



Acid and Base Reactions

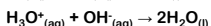
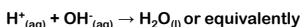
Acids and bases react with one another produce a salt and water (often called a neutralization reaction)



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What is the net ionic equation for this reaction????

The net ionic equation for any strong base / strong acid reaction is:

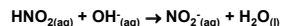


This will not be the case for weak acids and/or base reactions.

Ex.

Write a balanced net ionic equation for the reaction between nitrous acid and lithium hydroxide.

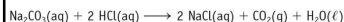
Answer:



Gas Forming Reactions

Table 5.3 Gas-Forming Reactions

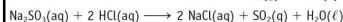
Metal carbonate or bicarbonate + acid \rightarrow metal salt + $\text{CO}_2(\text{g})$ + $\text{H}_2\text{O}(\ell)$



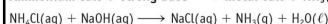
Metal sulfide + acid \rightarrow metal salt + $\text{H}_2\text{S}(\text{g})$



Metal sulfite + acid \rightarrow metal salt + $\text{SO}_2(\text{g})$ + $\text{H}_2\text{O}(\ell)$



Ammonium salt + strong base \rightarrow metal salt + $\text{NH}_3(\text{g})$ + $\text{H}_2\text{O}(\ell)$



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Oxidation-Reduction (Redox) Reactions:

Oxidation (Increase in oxidation number):

Gains Oxygen
Loses Hydrogen
Loses Electrons

Reduction (Decrease in oxidation number):

Loses Oxygen
Gains Hydrogen
Gains Electrons

Loses Electrons Oxidized; Gains Electrons Reduced
LEO goes GER

GER



Identification of Redox Reactions: Track changes in oxidation numbers of elements from reactants to products

Oxidation numbers: Except in the case of monoatomic ions, oxidation numbers are "virtual charges" for keeping track of how electrons are shared in a compound. They are assigned using a set of rules.

Assigning Oxidation Numbers (See page 200 in your text)

- In free elements (that is, in the uncombined state), each atom has an oxidation number of zero. Thus each atom in H_2 , Br_2 , Na , Be , K , O_2 , and P_4 has the same oxidation number: zero.
- For ions composed of only one atom, the oxidation number is equal to the charge on the ion. Thus Li^+ has an oxidation number of +1; Ba^{2+} ion, +2; Fe^{3+} , +3; I^- ion, -1; O^{2-} ion, -2, and so on. All alkali metals have an oxidation number of +1, and all alkaline earth metals have an oxidation number of +2 in their compounds. Aluminum has an oxidation number of +3 in all its compounds.
- The oxidation number of oxygen in most compounds (for example, MgO and H_2O) is -2, but in hydrogen peroxide (H_2O_2) and peroxide ion (O_2^{2-}), its oxidation number is -1.
- The oxidation number of hydrogen is +1, except when it is bonded to metals in binary compounds. In these cases (for example, LiH , NaH , and CaH_2), its oxidation number is -1.
- Fluorine has an oxidation number of -1 in all its compounds. Other halogens (Cl , Br , and I) have negative oxidation numbers when they occur as halide ions in their compounds. When combined with oxygen—for example in oxoacids and oxoanions, they have positive oxidation numbers.
- In a neutral molecule, the sum of the oxidation numbers of all the atoms must be zero. In a polyatomic ion, the sum of oxidation numbers of all the elements in the ion must be equal to the net charge of the ion. For example, in the ammonium ion, NH_4^+ , the oxidation number of N is -3 and that of H is +1. Thus the sum of the oxidation numbers is $-3 + 4(+1) = +1$, which is equal to the net charge of the ion.

Oxidation-Reduction (Redox) Reactions:

Oxidations and Reductions must occur in pairs (i.e. no oxidation without reduction and vice versa.)

The **oxidizing agent** oxidizes another substance (it itself is reduced) while the **reducing agent** reduces another substance (while being oxidized)

Remember: The agent refers to what the substance does to something else.

Even though only 1 element within a compound or ion may actually be oxidized or reduced, generally the whole compound or ion is referred to as the oxidizing or reducing agent.

Examples: Determine which of the following are redox reactions:

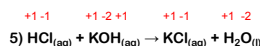
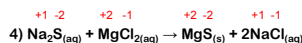
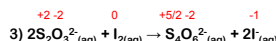
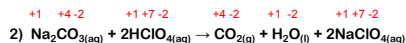
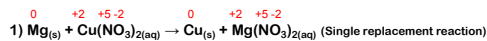
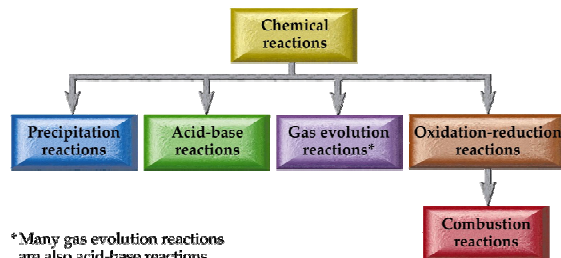


Table 5.4 Common Oxidizing and Reducing Agents

Oxidizing Agent	Reaction Product	Reducing Agent	Reaction Product
O ₂ , oxygen	O ²⁻ , oxide ion or O combined in H ₂ O	H ₂ , hydrogen	H ⁺ (aq), hydrogen ion or H combined in H ₂ O or other molecule
Halogen, F ₂ , Cl ₂ , Br ₂ , or I ₂	Halide ion, F ⁻ , Cl ⁻ , Br ⁻ , or I ⁻	M, metals such as Na, K, Fe, and Al	M ⁿ⁺ , metal ions such as Na ⁺ , K ⁺ , Fe ²⁺ or Fe ³⁺ , and Al ³⁺
HNO ₃ , nitric acid	Nitrogen oxides* such as NO and NO ₂	C, carbon (used to reduce metal oxides)	CO and CO ₂
Cr ₂ O ₇ ²⁻ , dichromate ion	Cr ³⁺ , chromium(III) ion (in acid solution)		
MnO ₄ ⁻ , permanganate ion	Mn ²⁺ , manganese(II) ion (in acid solution)		

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Summary of Common Reaction Types in Aqueous Solutions



*Many gas evolution reactions are also acid-base reactions.

Reactions that remove ions from solution, produce a solid precipitate, a water molecule or a gas molecule are product favored reactions. Those that don't occur to an appreciable amount are said to be reactant favored.

Solution Concentration: Molarity

$$\text{Molarity (M)} = \frac{\text{moles solute}}{\text{Liters of solution}}$$

Units = mol/L

Example:

65.00g of sodium phosphate (163.94g/mol) is dissolved in enough water to produce 650.mL of solution.

- Determine the molarity of the solution.
- Determine the number of sodium ions in a 50.0mL sample of this solution.

Answer:

a) $65.00\text{g} / 163.94\text{g/mol} = .396486 \dots$ mol sodium phosphate
 $.3965\text{mol} / .650\text{L} = .609979 \dots = \mathbf{.610M}$

b) Sodium phosphate = Na₃PO₄
 $.610\text{mol/L} (.0500\text{L})(6.022 \times 10^{23}\text{units/mol})(3\text{sodium ions/unit})$
 $= 5.50994 \dots \times 10^{22} = \mathbf{5.51 \times 10^{22} \text{Na}^+ \text{ions}}$

Solution Concentration: Dilutions

To dilute an existing solution use the relationship

$$M_1V_1 = M_2V_2$$

Where M is the molarities of the solutions and V is the volumes of the solutions (you can use mL since it is a ratio)

Example:

A student wants to produce 75.0mL of a .150M HCl solution from a stock solution that is 6.00M. Determine the procedure required to produce the desired solution.

Answer:

$$6.00M(V_1) = (.150M)(.0750\text{L})$$

$$V_1 = .001875\text{L} = 1.88\text{mL}$$

Use 1.88mL from the stock solution and fill with distilled water to the 75.0mL mark.

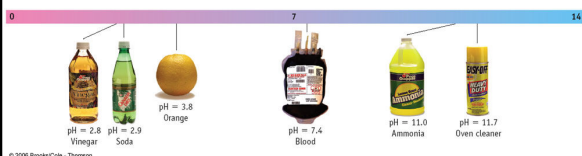
Solution Concentration: pH

For acids and bases pH determines the concentration of the hydrogen ion in solution (in molarity).

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$$

Brackets always means molarity. H_3O^+ and H^+ refer to the same thing.



Sample:

1) Determine the pH of a .025M perchloric acid solution.

Answer: Perchloric Acid (HClO_4) is a strong acid so 100% dissociation can be assumed.

.025M HClO_4 implies .025M in H^+ (H_3O^+) ions

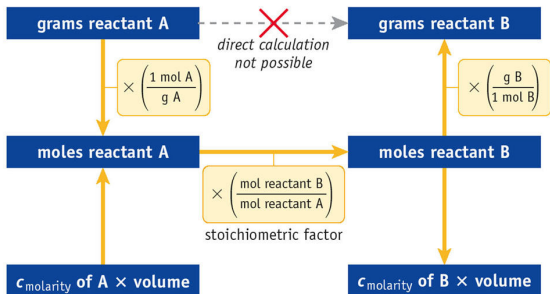
$$\text{pH} = -\log(.025) = 1.6020\dots = 1.6$$

2) Calculate the hydronium ion concentration for a solution with a pH of 8.3.

Answer:

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} = 10^{-8.3} = 5.01187\dots \times 10^{-9}\text{M} = 5.0 \times 10^{-9}\text{M}$$

Solution Stoichiometry:



Conversion Map

Example:

If you combine 75.0mL of 0.350M HCl and an excess of Na_2CO_3 , what mass of CO_2 , in grams, is produced?

Answer:

Balanced Equation:



$$\text{mol HCl} = (.0750\text{L})(0.350\text{M}) = .02625\text{mol HCl}$$

$$\text{mole ratio: } .02625 \text{ mol HCl} (1\text{mol CO}_2 / 2\text{mol HCl}) = .013125\text{mol CO}_2$$

$$\text{Mass CO}_2: .013125 \text{ mol CO}_2 (44.01\text{g/mol}) = .578 \text{ g CO}_2$$

Titrations:

Used for determining the unknown concentration/amount of a given substance by reacting it in a stoichiometric ratio with a solution of known concentration (the standard solution) or primary standard (pure solid acid or base of precisely known mass).

Definitions:

Titrant: The substance of known concentration. The substance you are "titrating" something with.

Equivalence Point: That point in the reaction when the exact amount of the titrant has been used to react completely with the unknown in the correct stoichiometric ratio.

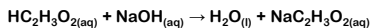
Acid Base Indicator: A substance added to a solution that changes color at a specific pH to identify the equivalence point.

Endpoint: That point in the reaction when the indicator changes color. Usually, but not always, corresponds to the equivalence point.

Standardization: To precisely determine the concentration of a solution.

Example:

1. A 25.0mL sample of vinegar (which contains the weak acid acetic acid) requires 28.33mL of a 0.953M solution of NaOH for titration to the equivalence point. What mass of acetic acid, in grams, is in the vinegar sample, and what is the concentration of acetic acid in the vinegar?



Answer:

$$.02833\text{L}(0.953\text{M}) = .0270\text{mol NaOH} = \text{moles of acetic acid (1:1)}$$

$$.0270\text{mol} (60.0520\text{g/mol}) = 1.62\text{g acetic acid}$$

$$\text{Molarity} = .0270\text{mol} / .0250\text{L} = 1.08\text{M}$$

Example:

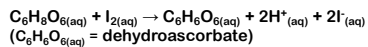
2. Hydrochloric acid can be purchased from chemical supply houses with a concentration of 0.100M, and this solution can be used to standardize the solution of a base. If titrating 25.00mL of a sodium hydroxide solution to the equivalence point requires 29.67mL of 0.100M HCl, what is the concentration of the base?

Answer:

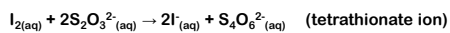
$$\text{Mol HCl} = (0.100\text{M})(.02967\text{L}) = .002967\text{mol HCl} = \text{mol NaOH (1:1)}$$
$$\text{Molarity NaOH} = .002967\text{mol NaOH} / .02500\text{L} = .119\text{M}$$

Redox Reaction Titration Sample Problem:

Vitamin C, ascorbic acid ($\text{C}_6\text{H}_8\text{O}_6$), is a reducing agent. One way to determine the ascorbic acid content of a sample is to mix the acid with an excess of iodine,



And then titrate the iodine that did not react with the ascorbic acid with sodium thiosulfate. The balanced, net ionic equation for the reaction occurring is this titration is



Suppose 50.00mL of 0.0520M I_2 was added to the sample containing ascorbic acid. After the ascorbic acid/ I_2 reaction was complete, the I_2 not used in the reaction required 20.30mL of 0.196M $\text{Na}_2\text{S}_2\text{O}_3$ for titration to the equivalence point. Calculate the mass of the ascorbic acid in the unknown sample.

$$\text{I}_2 \text{ originally} - (.0520 \text{ L})(.0500 \text{ M}) = .0026 \text{ mol}$$

I_2 titrated -

$$(.0203 \text{ L})(.196 \text{ M}) = .00398 \text{ mol S}_2\text{O}_3^{2-} \quad (1 \text{ mol I}_2 / 2 \text{ mol S}_2\text{O}_3^{2-})$$
$$= .00199 \text{ mol I}_2$$

I_2 reacted with $\text{C}_6\text{H}_8\text{O}_6$ -

$$.0026 \text{ mol} - .00199 \text{ mol} = .000611 \text{ mol I}_2$$

$\text{C}_6\text{H}_8\text{O}_6$ used -

$$.000611 \text{ mol I}_2 (1 \text{ mol C}_6\text{H}_8\text{O}_6 / 1 \text{ mol I}_2)(174.108 \text{ g/1 mol})$$
$$= .106 \text{ g C}_6\text{H}_8\text{O}_6$$