Solutions, Acids and Bases

Prentice Hall *Physical Science* –

Chapter 8

Review – What is a solution?

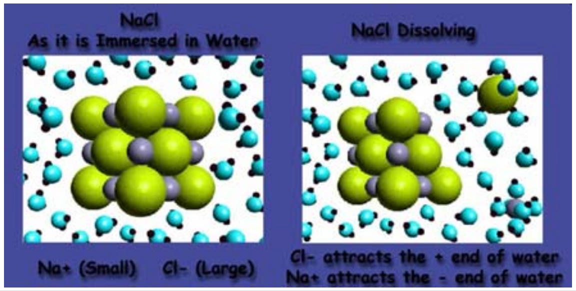
- *Solution* – homogeneous mixture formed when one substance dissolves into another
  - Any state of matter (solid, liquid, gas) can become part of a solution
  - Examples: see Table at top of p. 229
    - Air (O_2 and CO_2 dissolved in N_2)
    - Stainless Steel (Chromium and Nickel dissolved in Iron)
    - Sugar water
- *Solute* – substance whose particles are dissolved
- *Solvent* – substance in which the particles dissolve

Dissolving

- We will focus on water solutions (water is the solvent)
- Substances can dissolve in water in 3 ways:
  - Dissociation of ionic compounds
  - Dispersion of molecular compounds
  - Ionization of molecular compounds
- To dissolve, attractions that hold the solute together and the solvent together must be overcome:
  - Before dissolving, solvent particles are attracted to each other, and solute particles are attracted to each other…
  - Solute and solvent particles must attract one another

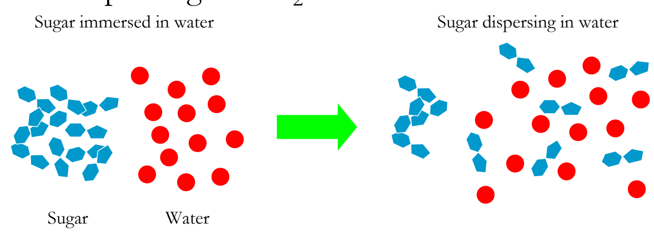
Dissociation of Ionic Compounds

- *Dissociation* – an ionic compound separates into ions as it dissolves
- Dissolution by dissociation is a *physical change*
- Example: NaCl in H_2O



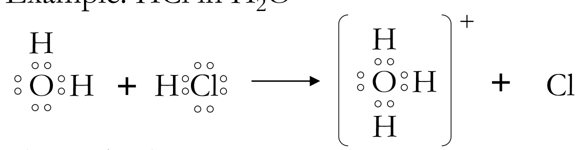
Dispersion of Molecular Compounds

- *Dispersion* – solute breaks into small pieces that spread throughout solvent
- Dissolution by dispersion is a physical change
- Example: sugar in H_2O



Ionization of Molecular Compounds

- *Ionization* – neutral molecules gain or lose electrons for form ions
- Dissolution by ionization is a *chemical change*: a reaction takes place between the solute and solvent
- Example: HCl in H_2O



(See p. 230 in text)

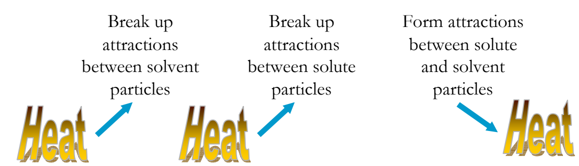
Properties of Liquid Solutions

- Properties of solutions differ from individual properties of solute and solvent
- Example: NaCl in H_2O

	Substance/Mixture			
Property	NaCl	H_2O	NaCl dissolved in H_2O	
Conductivity	Poor Conductor	Poor Conductor	Good Conductor	
Freezing Point	801°C	0°C	About -5°C to -15°C	← Freezing Point Depression
Boiling Point	1465°C	100°C	About 101°C to 105°C	← Boiling Point Elevation

Heat of Solution

- During formation of solutions, heat is either released (*exothermic*) or absorbed (*endothermic*)
- Examples:
  - NaOH (sodium hydroxide) in H_2O releases heat
  - NH_4NO_3 (ammonium nitrate) in H_2O absorbs heat (ammonium nitrate is used in cold packs)



Factors Affecting Rates of Dissolving

Think of adding sugar to your tea…

- Would you rather add a cube of sugar or granulated sugar? Why?
  - ***Surface Area*** – the greater the surface area of a solid solute, the more frequent the collisions between solvent and solute particles
  - *Large surface area = faster dissolution*
- Would you rather stir the sugar in or let it sit? Why?
  - ***Stirring*** – agitation moves dissolved particles away from the solid surface, allowing more solvent particles to contact the undissolved solid
  - *Stirring = faster dissolution*
- Does sugar dissolve faster in hot tea or iced tea? Why?
  - ***Temperature*** – increasing the temperature makes solvent particles move faster on average, increasing both the # of collisions and the energy of the collisions
  - *Increased temperature = faster dissolution*

Solubility

- ***Solubility*** – maximum amount of a solute that dissolved in a given amount of solvent at a constant temperature
- Usually given as g of solute per 100g of solvent
- Which substance given is most soluble in water? Which least?

Solubility in 100g of Water at 20°C	
Compound	Solubility (g)
NaCl	36.0
Baking Soda ($NaHCO_3$)	9.6
Sugar ($C_{12}H_{22}O_{11}$)	203.9

(Figure 8 p. 235 in text)

Solubility (cont'd)

- Solutions are described as saturated, unsaturated or supersaturated depending on how much solute is dissolved
- ***Saturated*** – contains as much solute as the solvent can hold (solvent is "filled")
- ***Unsaturated*** – contains less than the maximum amount of solute
- ***Supersaturated*** – contains more solute than can normally be held (solvent is "overloaded" – very unstable!)

Factors Affecting Solubility

- ***Solvent Polarity***
  - "Like dissolves like" – non-polar substances (like oil) are not soluble in polar substances (like vinegar), and vice-versa
  - Soap helps oil dissolve in water because soap has a non-polar end and a polar end (see Figure 11, p. 237 in text)
- ***Temperature***
  - Solubility of solids *increases* as temperature increases
  - Solubility of gases *decreases* as temperature increases
- ***Pressure***
  - Solubility of gases *increases* as pressure increases

Acids

- *Definition 1*: Produce hydronium ions (H_3O^+) in water
- Properties of Acids:
  - Sour taste (citrus – citric acid, vinegar – acetic acid, spoiled mile – butyric acid)
  - React with metals (remember the H2 balloon!)
- Common acids
  - Acetic acid (vinegar)
  - Carbonic acid (carbonated beverages)
  - Hydrochloric acid (stomach acid)
  - Sulfuric acid (car batteries)

Bases

- *Definition 1*: Produces hydroxide ions (OH^-) in water
- Properties of Bases
  - Bitter taste (unsweetened chocolate – theobromine)
  - Slippery feel (many soaps are made with bases)
- Common Bases
  - Aluminum hydroxide (deodorant, antacid)
  - Calcium hydroxide (concrete, plaster)
  - Magnesium hydroxide (antacid, laxative)
  - Sodium hydroxide (drain cleaner, soap production)

Neutralization and Salts

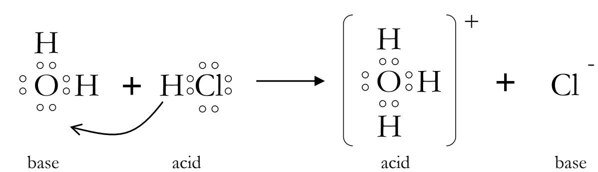
- *Neutralization*: reaction between acid and base
- What are the *reactants* in a neutralization reaction?
  - Acid
  - Base
- What are the *products* in a neutralization reaction?
  - Salt
  - Water

$$\underset{\textbf{acid}}{(H_3O^+ + Cl^-)} + \underset{\textbf{base}}{(Na^+ + OH^-)} \rightarrow \underset{\textbf{water}}{2H_2O} + \underset{\textbf{salt}}{(Na^+ + Cl^-)}$$

$$\underset{\textbf{acid}}{(H_3O^+ + NO_3^-)} + \underset{\textbf{base}}{(NH_3)} \rightarrow \underset{\textbf{water}}{H_2O} + \underset{\textbf{salt}}{(NH_4^+ + NO_3^-)}$$

Proton Donors and Acceptors

- *Recall that H has one proton and one electron. If H loses its one e^-, it is known as a hydrogen ion, H^+, and is just one proton.*
- Acids and bases can also be defined based on whether they lose (donate) or gain (accept) protons.
  - *Acid definition 2*: Proton donor
  - *Base definition 2*: Proton acceptor



pH Scale

- pH is a measure of hydronium ion concentration (acid concentration)
- Low pH = High H_3O^+ concentration (high acidity)
- pH scale goes from 0 to 14
  - 0 – 6.9 = ACID
  - 7 = NEUTRAL
  - 7.1 – 14 = BASE

Complete Reactions vs. Equilibrium (See section 7.5 in text)

- *Complete reaction*: all reactants are converted to products
- *Equilibrium*: state in which forward and reverse change happen at the same rate
- *Physical Equilibrium*: a physical change does not go to completion (like water vapor above liquid water)
- *Chemical Equilibrium*: chemical reaction does not go to completion (will have a mixture of products and reactants)
- Chemical equilibrium is represented by a reaction with both FORWARD and REVERSE arrows in the reaction equation.

Strong Acids and Bases

- Recall that acids and bases dissolve by *ionization* – a chemical change
- As certain acids and bases dissolve, the formation of ions goes to completion. These are STRONG ACIDS AND BASES.
- Examples:
  - HCl (hydrochloric acid)
  - NaOH (sodium hydroxide)

Weak Acids and Bases

- Weak acids and bases only ionize *slightly* in water – they form a chemical equilibrium in solution.
  - Weak acid: acetic acid

    $$CH_3COOH + H_2O \rightleftharpoons CH_3COO^- + H_3O^+$$
  - Weak base: ammonia

    $$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$
- Buffers: solutions that resist large changes in pH
  - Made using either a weak acid and its salt, or a weak base and its salt
  - Many biological applications

Electrolytes

- Electrolyte: substance that ionizes or dissociates into ions when dissolved in water
- Solutions of electrolytes are conductive
- Strong acids, strong bases, and soluble salts are strong electrolytes
- Examples of electrolytic solutions:
  - Sports drinks
  - Battery fluids