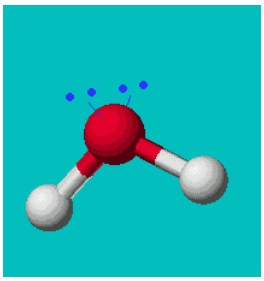
CHEMISTRY Ch. 13 Notes: Water and Its Solutions

NOTE: Vocabulary terms are in **boldface and underlined**. Supporting details are in *italics.*

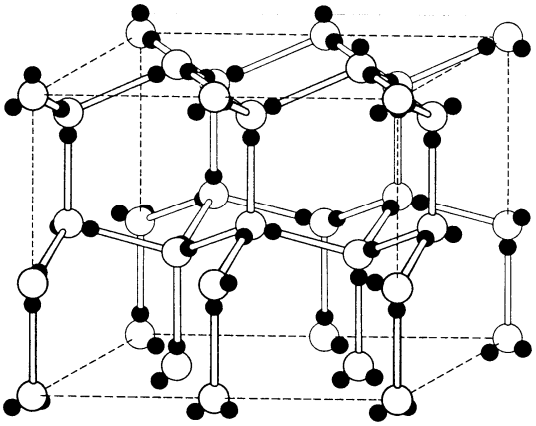
13.1 Notes

I. Water Molecule Characteristics



POLAR molecule (a dipole)

HYDROGEN BONDING in water



ICE CRYSTAL

A) colorless and odorless
B) neutral pH of 7
C) *triatomic* (three atoms)
D) *angular* shape with two unshared electron pairs
E) *polar (δ+ and δ- areas)*
F) **hydrogen bonds**—*an attraction between hydrogen and an unshared pair of an electronegative element on a neighboring molecule*
G) **surface tension**—*attraction between molecules on the surface of a liquid*
    1) surface tension makes water bead
    2) **surfactants**—*"wetting agents" which decrease surface tension* by breaking hydrogen bonds (soaps)
H) **capillarity (capillary action)—***moving upward, against gravity* (up through roots, etc.)
I) hexagonal crystals
J) *high specific heat*: 4.184 J/g°C
K) *high boiling point*: 100 °C

II. Heat Capacity of Water

A) **heat capacity**—amount of heat required to change a substance's temperature by exactly 1 °C

B) **specific heat capacity** *(specific heat) (C or c_p)—amount of heat required to raise the temperature of 1 g of a substance by exactly 1 °C*

C) *water has a high specific heat (4.184 J/g°C) as compared to other substances*:

air: 1.01	stainless steel: 0.51	aluminum: 0.90	gold: 0.13	wood: 1.76

D) *high requirements and retention = heat up slowly and cool down slowly*

III. Phase changes of evaporation and freezing

A) evaporation
1) water absorbs a large amount of heat before it changes phase from a liquid to a vapor
2) many hydrogen bonds must be broken
3) liquid has a lower K.E.; vapor has a higher K.E.

B) freezing
1) *solid water is less dense than liquid water* (ice floats)
2) *at 4 °C, water is densest*
3) *below 4 °C, it expands to form hexagonal crystals*
4) density comparisons
a) pure water before boiling: 0.9584 g/cm^3
b) pure water at 4 °C: 1.00 g/cm^3
c) density of pure ice: 0.9168 g/cm^3

13.2 Notes

IV. Aqueous Solutions

A) **aqueous solutions** *(aq)—water containing dissolved materials*
1) solutions are *homogeneous mixtures*
2) true solutions will not precipitate (form solids) or separate into layers on their own

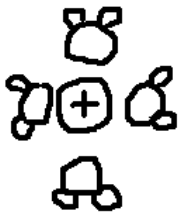
B) parts of a solution
1) **solute**—*the substance being dissolved*
2) **solvent**—*the substance doing the dissolving*

C) concentration
1) **dilute**—more solvent than solute (weak)
2) **concentrated**—more solute than solvent (strong)

D) **solvation**—*the dissolving of a solute into a solvent*

E) *hydration—a specific kind of salvation, when water is the solvent*

water hydrating an anion



water hydrating a cation

V. Electrolytes and Nonelectrolytes

A) **electrolytes**—*conductors in aqueous solution* (usually ionic compounds and acids)
1) **weak electrolyte**—*a small fraction of the solute is as free ions*
2) **strong electrolyte—***almost all of the solute is as free ions*

B) **nonelectrolytes**—*nonconductors in aqueous solution* (usually molecular and organic compounds)

VI. Suspensions and Colloids

A) suspension

B) **colloids** (also called **colloidal suspensions)**

1) *heterogeneous mixtures with two phases of intermediate particle sizes*
2) cannot be filtered or settled
3) colloids show **Brownian motion**
4) Examples: liquid aerosols, solid aerosols, emulsions, sols, pastes, gels, foam

PARTICLE SIZE:	*solution* <	*colloid (colloidal suspension)*	<	*suspension*
	salt water;	whipped cream;		dirt in water;
	acids; "air"	mayo; milk		wax in water

C. **Tyndall effect**—visible light transmitted by scattering through a colloid or suspension

1) light will not show a path thorough a solution
2) light will show a path through a colloid and a suspension

D. **emulsions**

1) *colloidal dispersions of liquid in liquid*
2) require an "emulsifier" like soap

VII. Water of Hydration (review)

A) **water of hydration**—*water molecules chemically integrated into a crystalline structure*

B) *hydrate—a compound with water in its structure*

1) *general formula: [compound]· H_2O*
2) *naming: [compound name] (prefix)hydrate*

$CuSO_4 \cdot 5H_2O$ copper(II) sulfate pentahydrate

C) *anhydrous—a hydrate without its water of hydration*

$Na_2SO_4 \cdot 2H_2O \rightarrow Na_2SO_4 + 2H_2O$

sodium sulfate dihydrate anhydrous sodium sulfate + water

D) **efflorescence**—the release of water by a hydrate (heating not needed)

E) **hygroscopic**—remove water from the atmosphere

F) **deliquescence**—absorbing excess water from the atmosphere to form a liquid substance

G) **desiccant**

1) a *drying agent* which is hygroscopic
2) examples: Damp Rid; packets of silica powder in shoe boxes

VIII. Solution Formation (abbreviation for solution = *soln.*)

A) *SOLUTE + SOLVENT = SOLUTION*

B) "**Like dissolves like**"

SOLUTE		**SOLVENT**	**SOLUTION?**
Polar	**+**	**Polar**	**Yes**
Polar	**+**	**Nonpolar**	**No**
Nonpolar	**+**	**Polar**	**No**
Nonpolar	**+**	**Nonpolar**	**Yes**

POLAR / IONIC examples:		
All BI	**Salts**	**Water**
All TI	**Sugars**	**Non-symmetrical molecules**
All crisscrossed formulas	**Acids (vinegar, etc.)**	

NONPOLAR EXAMPLES:		
Oil	**Wax/Parrafin**	**Symmetrical molecules**
Hydrocarbons: methane, propane, butane, octane…		

C) Factors which increase solution (dissolving) rate:
1) crush up solids (increase surface area [S.A.])
2) stirring
3) heating

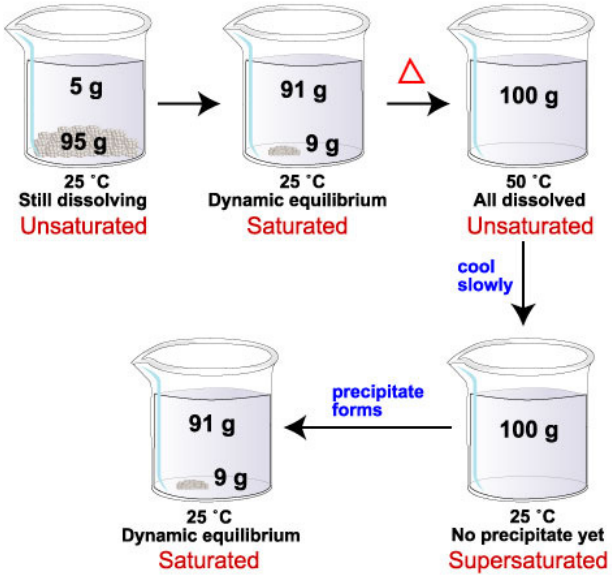
D. examples of different types of solutions
solute-solvent:
1) liquid-liquid: alcohol in water
2) solid-liquid: sugar in water
3) solid-solid: metal alloys such as brass
4) gas-gas: atmospheric air

IX. **Solubility**

A) *the maximum amount of a solute* (substance being dissolved) *which will dissolve in a given amount of solvent* (substance doing the dissolving)

B) **miscible**—*the ability of liquids to mix* (*opposite* = **immiscible**)

C) solution descriptions
1) **unsaturated**—small amount of solute; completely dissolved; room for more
2) **saturated**—too much solute; not all dissolved; excess settles on the bottom
3) **supersaturated**—an unstable solution formed from heating a saturated solution until all the solid dissolves; can recrystallize when cooled



X. **Molarity**

A) **Molarity (M)** = *moles of solute / liters of soln.*

B) **concentration [ ]** – *amount of solute / amount of solvent*

1) **dilute**—*weak* solution with little dissolved solute
2) **concentrated**—*strong* solution with large amounts of dissolved solute

C) examples

E1) What is the molarity of a solution of 0.6784 mol NaCl in 4.5 L water?

$$M = \frac{\text{moles}}{\text{L}} \qquad \frac{0.6784 \text{ mol NaCl}}{4.5 \text{ L soln.}} = 0.15 \frac{\text{mol}}{\text{L}} = \boxed{0.15 \text{ M}}$$

E2) A sugar solution contains 12.5 g of $C_{12}H_{22}O_{11}$ dissolved in 500.00 mL of water. What is the molarity of the solution?

$$M = \frac{\text{moles}}{\text{L}} \qquad 12.5 \text{ g } C_{12}H_{22}O_{11} \times \frac{1 \text{ mol } C_{12}H_{22}O_{11}}{342.0 \text{ g } C_{12}H_{22}O_{11}} = 0.0365 \text{ mol } C_{12}H_{22}O_{11}$$

$$500.00 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.50000 \text{ L} \qquad \frac{0.0365 \text{ mol}}{0.50000 \text{ L}} = \boxed{0.0730 \text{ M}}$$

E3) How many grams of KBr should be added to 977.6 mL of water to make a 3.0 M solution?

$$977.6 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.9776 \text{ L}$$

$$3.0 \frac{\text{mol KBr}}{\text{L}} \times 0.9776 \text{ L} \times \frac{119.0 \text{ g KBr}}{1 \text{ mol KBr}} = \boxed{350 \text{ g KBr}}$$

XI. **Colligative Properties** of Solutions

A) these properties relate to the number of solute particles

B) examples

1) vapor pressure elevation (volatile solute α v.p.)
2) **boiling point elevation** (BPE α solute concentration)
3) **freezing point depression** (FPD α solute concentration)