Chemistry 6A F2007

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Heat and the Specific Heat Capacity:

✧ When heat is absorbed or lost by a body, the temperature must change as long as the phase (*s, g or l*) remains constant.

✧ The amount of heat (q) transfer is related to the mass and temperature by:

q = heat lost or gained (J) m = mass of substance (g)

$$q = m \times C \times \Delta T$$

C = the Specific Heat Capacity of a compound $\left(\frac{J}{g \cdot {}^{o}C}\right)$

ΔT is the temperature change in degrees Celsius or Kelvins

Changes of State:

Energy Changes Accompanying Phase Changes

The sequence:

heat solid → form liquid → heat liquid → form gas (vapor) → heat gas

↑ *(melt)* ↑ *(boil or vaporize)*

is ***endothermic***. *(energy must be supplied)*

The sequence:

cool gas → form liquid → cool liquid → form solid → cool solid

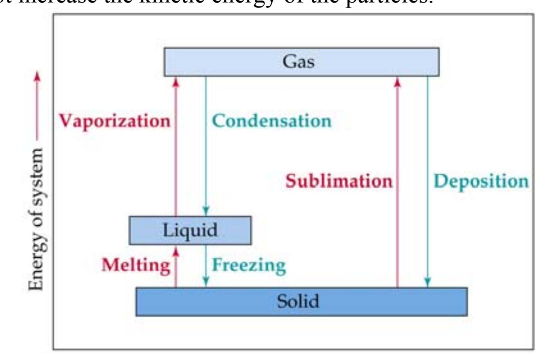
↑ *(condense)* ↑ *(freeze)*

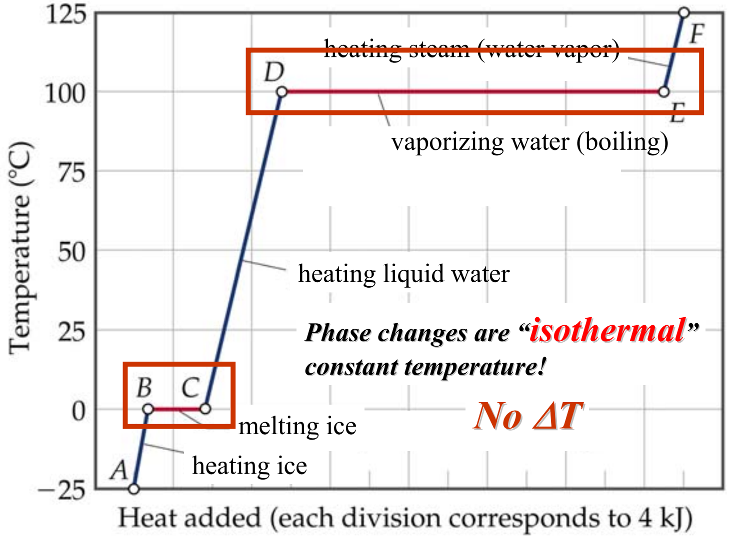
is ***exothermic***. *(energy is released)*

Changes of State:

All phase changes are ***isothermal***, that is a phase change occurs at one temperature.

The heat added or lost involves overcoming the intermolecular forces. It does not increase the kinetic energy of the particles.





How much energy is needed to take 10.0g of ice from –10.0°C to wateer at 10.0°C.

The total energy required is a sum of the steps which take the ice to water.

$$\underset{(-10.0^{o}C)}{\text{Ice } H_2O(s)} \xrightarrow[]{\textit{use } C_{ice}} \underset{(0^{o}C)}{\text{Ice } H_2O(s)} \xrightarrow[]{\textit{use } Heat_{(fus)}} \underset{(0^{o}C)}{\text{water } H_2O(l)} \xrightarrow[]{\textit{use } C_{water}} \underset{(10.0^{o}C)}{\text{water } H_2O(l)}$$

C_{ice}: $\frac{2.092J}{g\cdot^{o}C}$ C_{water}: $\frac{4.184J}{g\cdot^{o}C}$ $Heat_{(fus)}$: $\frac{333J}{g}$

Where do you get these numbers…

Back
of (the)
Book

How much energy is needed to take 10.0g of ice from –10.0°C to wateer at 10.0°C.

$$\text{Heat} = m \times C \times \Delta T \qquad \text{Heat} = m \times \text{Heat}_{(fus)}$$

Total Heat:

$$= 10.0\text{g ice} \times \frac{2.092J}{g^{o}C} \times \{0^{o}C - (-10.0^{o}C)\}$$ *warming the ice*

$$+10.0g \times \frac{333J}{g}$$ *melting the ice*

$$+10.0g \times \frac{4.184J}{g^{o}C} \times (10.0^{o}C - 0.0^{o}C)$$ *warming the water*

= 3960 J or 946 cal

Chapter 7: *Solutions and Colloids*

LEARNING OBJECTIVES

1. Classify mixtures and identify the components.
2. Predict solubility using solute-solvent interactions and understand the solution process.
3. Understand how to prepare solutions and calculate concentrations in units of molarity, weight/weight percent, weight/volume percent, and volume/volume percent.
4. Perform stoichiometric calculations using solution concentration and volume.
5. Calculate boiling point elevation, freezing point depression and the osmotic pressure of aqueous solutions.
6. Describe the characteristics of colloids.
7. Describe the process of dialysis and contrast it to the process of osmosis.

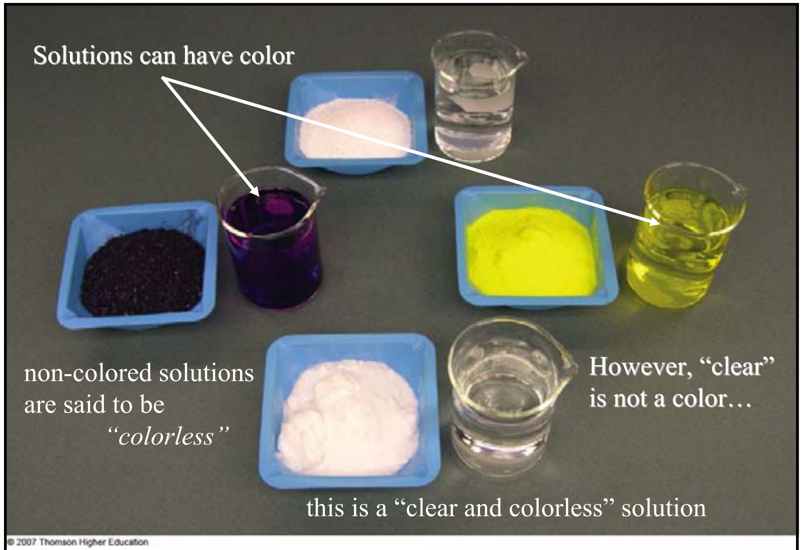
Solubility and the Solution Process:

When a *solute* dissolves in a *solvent*, the *solution* that forms is a homogeneous mixture.

solute: “That which is dissolved” (generally the lesser quantity)

solvent: “That which is dissolved” (generally the greater quantity)

- Solutions can have more than one solutes
- Solutions can be liquid or gaseous
- Solutions with water as the solvent are called “*aqueous*” solutions
- The particles dissolved in the solvent are too small to reflect light so all solutions are “clear”
- The particles dissolved in the solution never settle out due to gravity.



SOLUBILITY:

The solubility of a solute is the maximum amount of solute that can be dissolved in a specific amount of solvent under specific conditions of temperature and pressure.

General Solubility Guidelines

TABLE 7.3 Approximate solubility terms

Solute solubility (g solute/100 g H_2O)	Solubility term
Less than 0.1	Insoluble
0.1–1	Slightly soluble
1–10	Soluble
Greater than 10	Very soluble

TABLE 7.2 Examples of solute solubilities in water (0°C)

Solute: Name	Formula	Solubility (g solute/100 g H_2O)
Ammonium chloride	NH_4Cl	29.7
Ammonium nitrate	NH_4NO_3	118.3
Ammonium orthophosphate	$NH_4H_2PO_4$	22.7
Ammonium sulfate	$(NH_4)_2SO_4$	70.6
Calcium carbonate	$CaCO_3$	0.0012
Calcium chloride	$CaCl_2$	53.3
Calcium sulfate	$CaSO_4$	0.23
Potassium carbonate	K_2CO_3	101
Potassium chloride	KCl	29.2

Notice that ammonium and potassium salts are quite soluble…

TABLE 7.2 Examples of solute solubilities in water (0°C)

Solute: Name	Solute: Formula	Solubility (g solute/100 g H_2O)
Sodium bicarbonate	$NaHCO_3$	6.9
Sodium bromide	NaBr	111
Sodium carbonate	Na_2CO_3	7.1
Sodium chloride	NaCl	35.7
Sodium iodide	NaI	144.6
Ascorbic acid (vitamin C)	$C_6H_8O_6$	33
Ethyl alcohol	C_2H_5OH	∞[a]
Ethylene glycol (antifreeze)	$C_2H_4(OH)_2$	∞
Glycerin	$C_3H_5(OH)_3$	∞
Sucrose (table sugar)	$C_{12}H_{22}O_{11}$	179.2

Miscible liquids are infinitely soluble in one another!

A SATURATED SOLUTION is a stable solution that contains the maximum amount of dissolved solute under the normal conditions of temperature and pressure.

A SUPERSATURATED SOLUTION is an unstable solution that contains an amount of solute greater than the solute solubility under the prevailing conditions of temperature and pressure.

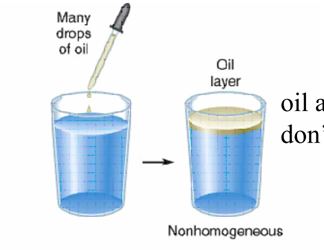
The solute in a supersaturated solution will fall out if it is disturbed:

Solids that do not dissolve in a particular liquid solvent are called *precipitates*.

lead (II) iodide is an insoluble salt

and insoluble salt is a precipitate

Liquids that do not dissolve in a particular liquid are said to be *immiscible*.



oil and water don't mix!

INCREASING THE RATE OF DISSOLVING

Crush or grind the solute:

Smaller particles provide for more surface area for solvent interaction, thus increasing the rate of solubility.

Heat the solvent:

When the solvent molecules move faster, there are more frequent collisions with solute thus increasing the rate of solution.

Stir or agitate the solution:

Stirring removes locally saturated solution from the vicinity of the solute thus allowing unsaturated solvent to take its place.