Phases of Matter

There are three (3) major phases of matter that we will concern ourselves with in CP Chemistry.

Solid- has defined shape and defined volume. The atoms in a solid are tightly packed and while still moving really only vibrate in place.

Liquid- has defined volume but not shape. It will take the shape of its container. The atoms in a liquid are less tightly packed than in a solid allowing the atoms more freedom of movement.

Gas- has no defined volume or shape. It will take the shape and volume of its container. The atoms in a gas are very far apart, making gas phase mostly empty space. This allows the most amount of movement of atoms and allows gases to be compressed.

Changes in Phase of Matter

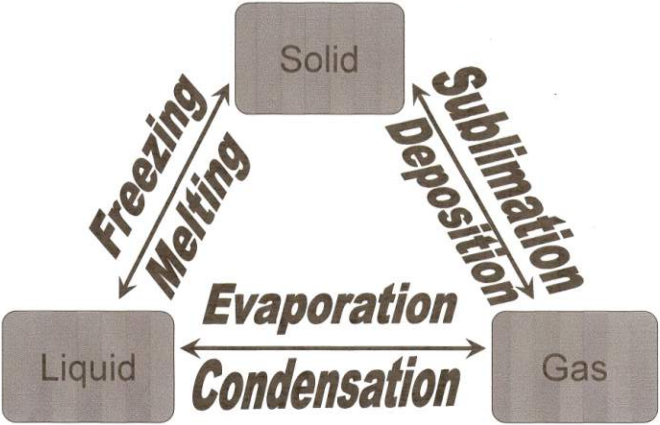
Phase changes are physical changes.

Solid to Liquid is melting
Liquid to Solid is freezing
Liquid to Gas is evaporating
Gas to Liquid is condensing
Gas to Solid is depositing
Solid to Gas is subliming

Note: The temperature at which these changes occurs depends on the substance that you are dealing with!

See Phase Change Diagram (triangle) for a more visual approach.

Phase Change Diagram



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Properties of Matter

<u>Physical Properties</u>- are characteristics of matter that can be discovered without changing the identity of the substance.

Examples: color, size, shape, state of matter, boiling point, freezing point, melting point

<u>Chemical Properties</u>- are characteristics of matter that can only be discovered by changing or attempting to change the identity of the substance.

Examples: reactivity, flammability, digestibility

Changes in Matter

<u>Physical Changes</u>- changes that do NOT result in the formation of a new substance or substances.

Examples: melting, freezing, evaporating, condensing, subliming, depositing, cutting, smashing, crushing, boiling.

<u>Chemical Changes</u>- changes that result in the formation of a new substance or substances.

Examples: burning, combustion, digestion, rotting, rusting, tarnishing, biological growth

<u>Law of Conservation of Matter</u>- In any change, matter can NOT be created or destroyed, only rearranged.

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Types or Classification of Matter

Pure Substances- types of matter that have a fixed composition.

Elements- have only one (1) type of atom

Examples: Hydrogen, Helium, Calcium, Sodium
Or any other from the periodic table.

Compounds- have two or more types of atoms (elements) combined in a fixed ratio.

Examples: Carbon Dioxide, Salt, Carbon Monoxide, Water

Mixtures- types of matter that have a variable composition.

Homogeneous- have the same composition throughout the sample

Solutions- all solutions are homogeneous mixtures.

Examples:

Solid solutions- brass, steel alloys
Liquid solutions- salt water, ice tea, coffee
Gas solution- air

Heterogeneous- have variation in composition throughout the sample.

Examples: Chocolate chip cookies, salsa, muddy water, milk,

These two usually refer to heterogeneous liquids

Suspensions- will settle out into layers over time. Muddy water and non-homogenized milk are two examples of suspensions.

Colloids- smaller particles than a suspension so it will not settle out over time. Homogenized milk and fog are two examples of colloids.

The Tyndall Effect can be used to determine if a liquid or gas sample is homogeneous or heterogeneous. A beam of light is passed through the sample (flashlight is sufficient). If the light is scattered the sample is a heterogeneous mixture, if it is not scattered the sample is a homogeneous mixture.

It is the Tyndall Effect that makes it necessary to use low beams and or fog lights rather than high beams when driving in foggy conditions.

Separation of Mixtures

Distillation- separation of a mixture using boiling points.

Filtration- separation of a mixture using particle size.

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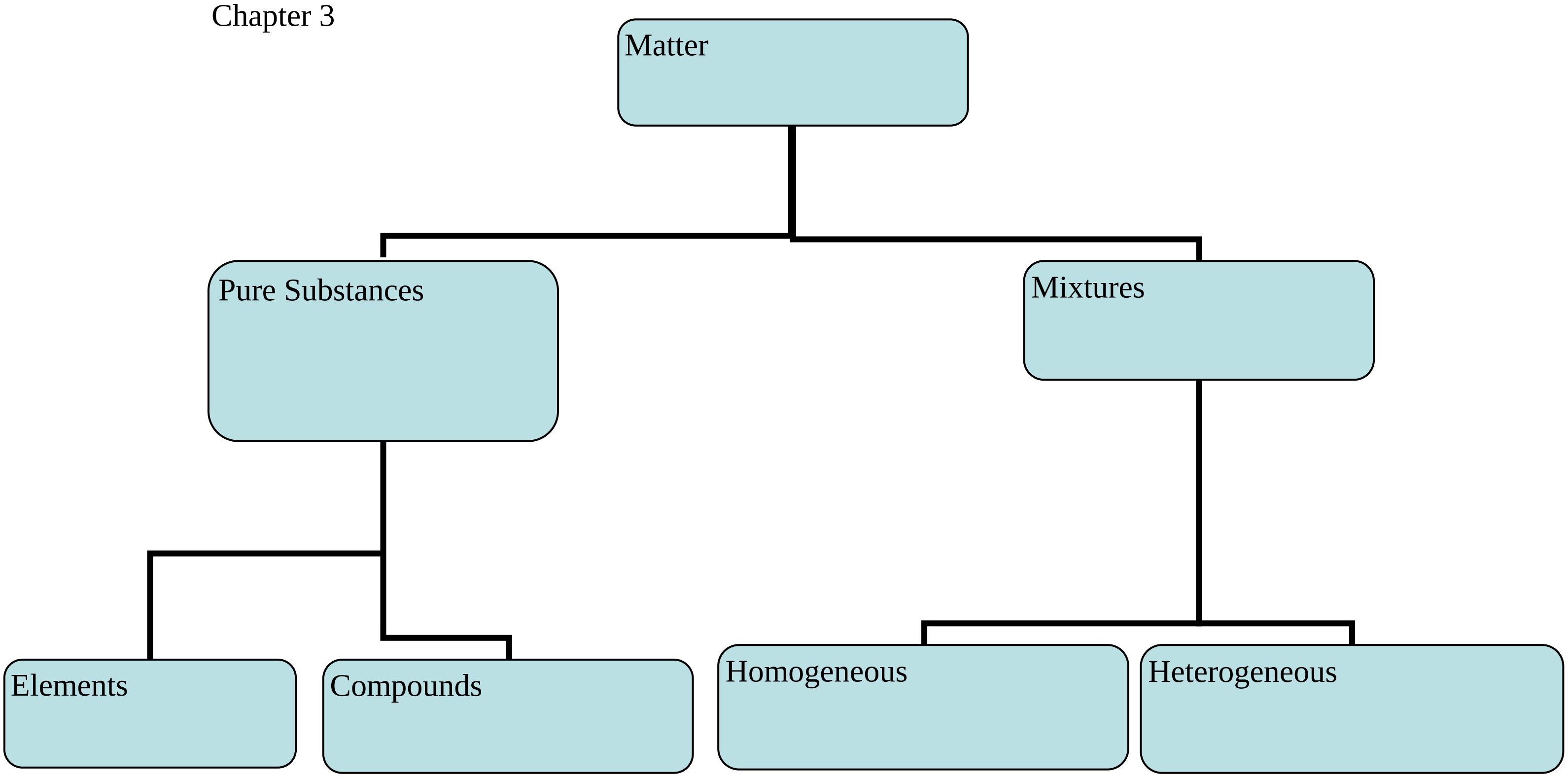

Energy and Energy Calculations

Energy: the ability to do work

Temperature: a measurement of the random motion of the components of a substance

Heat: flow of energy due to a temperature difference

Exothermic- release energy, energy exits the system

Endothermic- absorb energy, energy enters the system

calorie- amount of energy needed to change the temperature of 1 gram of water by 1°C

Calorie- nutritional calorie, actually 1000 calories or 1 kilocalorie

Joule- S.I. unit for energy

1 calorie = 4.184 Joules

Both calories and joules use the metric prefixes, you can have a kilocalorie or a kilojoule

You must be able to perform conversions with energy units.

calories to joules:

$$\frac{12 \text{ caloroies}}{} \left| \frac{4.184 \text{ joules}}{1 \text{ calorie}} \right. = 50.208 \text{ joules}$$ (proper sig. figs = 50. joules)

joules to calories:

$$\frac{455 \text{ joules}}{} \left| \frac{1 \text{ calorie}}{4.184 \text{ joules}} \right. = 108.747$$ (proper sig. figs = 109 calories)

calories to kilojoules:

$$\frac{1256 \text{ caloroies}}{} \left| \frac{4.184 \text{ joules}}{1 \text{ calorie}} \right. = 5255.104 \text{joules}$$

$$\frac{5255.104 \text{ joules}}{} \left| \frac{1 \text{ kilojoule}}{1 \text{ x}10^3 \text{ joules}} \right. = 5.255104 \text{ kilojoules}$$ (sig. figs = 5.255kJ)

joules to kilocalories:

$$1256 \text{ joules} \times \frac{1 \text{ calorie}}{4.184 \text{ joules}} = 300.1912 \text{ calories}$$

$$300.1912 \text{ calories} \times \frac{1 \text{ kilocalorie}}{1 \times 10^3 \text{ calories}} = 0.3001912 \text{ kilocalories}$$ (sig. figs = 0.3002 kcal)

kilocalories to joules:

$$1.04 \text{ kilocalories} \times \frac{1 \times 10^3 \text{ calorie}}{1 \text{ kilocalorie}} = 1040 \text{ calories}$$

$$1040 \text{ calories} \times \frac{4.184 \text{ joules}}{1 \text{ calorie}} = 4351.36 \text{ joules}$$ (sig. figs = 4350 joules)

kilojoules to calories:

$$486 \text{ kilojoules} \times \frac{1 \times 10^3 \text{ joules}}{1 \text{ kilojoule}} = 486000 \text{ joules}$$

$$486000 \text{ joules} \times \frac{1\text{calorie}}{4.184 \text{ joules}} = 116{,}156.7878.\text{calories}$$ (sig. figs = 116,000 calories)

Specific Heat- amount of energy needed to change the temperature of 1 gram of a substance by 1°C

Unit for specific heat = J/g°C (joules per gram degree Celsius)

Specific heat is used in energy calculations, Formula is $Q = S \times M \times \Delta T$, where

Q = energy in Joules
S = specific heat of the substance in J/g°C
M = mass of the sample in grams
ΔT = change in temperature

These calculations become algebra problems, I can ask you to solve for any component of the equation if I give you the other three.

Example 1. How much energy is needed to raise the temperature of 25 grams of aluminum (specific heat = 0.89 J/g°C) from 20.°C to 25°C?

I know S, M and ΔT so I will be solving for Q.

$Q = 0.89 \times 25 \times (25-20)$
$Q = 0.89 \times 25 \times 5$
Q= 111.25 joules (proper sig figs = 110 joules)

Example 2. If I apply 125 joules of energy to 25 grams of aluminum (specific heat = 0.89 J/g°C) by how many degrees will my temperature change?

I know Q, S and M so I will be solving for ΔT.

125 = 0.89 x 25 x ΔT
125 =22.25 x ΔT
ΔT =125/22.25
ΔT= 5.617977 °C (proper sig figs = 5.6°C)

Example 3. If I apply 125 joules of energy to a sample of aluminum (specific heat = 0.89 J/g°C) and my temperature changes from 30.°C to 80.°C, what is the mass of my sample?

I know Q, S and ΔT so I will be solving for M.

125 = 0.89 x M x (80-30)
125 =0.89 x M x 50
125 = 44.5 x M
M = 125/44.5
M = 2.8089 grams (proper sig figs = 2.8 grams)

Example 4. If I apply 130. joules of energy to a 100.0 gram sample and my temperature changes from 30.°C to 40.°C, what is the identity of my sample?

I know Q, M and ΔT so I will be solving for S. I can then use S and the chart on page 70 to determine the identity of the substance from the specific heat that I calculate.

130 = S x 100 x (40-30)
130 =S x 100 x 10
130 = S x 1000
S = 130/1000
S = 0.13 J/g°C (proper sig figs = 0.13 J/g°C)

I then look in the textbook on page 70, I find that gold has a specific heat of 0.13 J/g°C, therefore my sample must be gold.

Chapter 13 – Liquids and Solids

Section 13.1

Most substances become denser as they go from gas to liquid to solid.

Ice is an exception: Ice is less dense than liquid water.

Water expands as it freezes- breaks pipes, engine blocks

Ice floats on water , insulating ponds, lakes and other bodies of water which prevents them from freezing solid in winter which provides liquid water for aquatic life year round.

Focus on Water and its Phase Changes

Water covers 70% of the Earth

Forms a solid at 0°C- this is known as the **Normal Freezing Point**

Forms a gas at 100°C- this is known as the **Normal Boiling Point**

When we heat water (add energy) the molecules move faster. Eventually they will be moving fast enough to begin to form bubbles of gas (this happens when the temperature reaches 100°C). When the boiling point is reached, the temperature stops climbing. The temperature will remain at 100°C until all water is turned to vapor. Once all of the liquid water is vaporized, then the temperature will continue to rise.

This process can be illustrated as a heating/cooling curve (shown on page 401 of the text).

Section 13.2

REMEMBER: Changes in state of matter are physical changes, they do not chemically change the substance.

Intermolecular forces – energy that holds one water molecule to another water molecule

Intramolecular forces– energy that holds a water molecule together.

Heat of Fusion- amount of energy needed to melt solid to liquid

Heat of Vaporization- amount of energy needed to change liquid to vapor

EXAMPLE PROBLEMS:

Problem #1 : Calculate the energy required to melt 8.5 grams of ice at 0°C. The heat of fusion for ice is 6.02 kJ/18 g.

8.5 grams	6.02 kJ	=	2.842777778 kJ
	18 grams		

rounded to proper significant figures= 2.8 kJ

Problem #2 : Calculate the energy in kilojoules required to heat 25 grams of liquid water from 25°C to 100.0°C and change it to steam at 100.0°C. The specific heat capacity of water is 4.184 J/g°C and the heat of vaporization of water is 40.6 kJ/18 g.

This is a multi-step problem.

Step 1. Heat water from 25°C to 100.0°C.

Q= S x M x ΔT

Q= (4.184)(25)(75)

Q= 7845 J

I want my answer in kilojoules, so I might as well convert this step's answer now:

7845 J	1 kJ	=	7.845 kJ
	$1x10^3$ J		

Step 2. Change liquid at 100.0°C to gas at 100.0°C

25 grams	40.6 kJ	=	56.38888889 kJ
	18 grams		

Step 3. Add energy from each step to get total energy needed

7.845 + 56.38888889 = 64.23389 rounded to correct significant figures= 64 kJ

Problem #3 : Calculate the total energy required (in kilojoules) to melt 15 grams of ice at 0.0°C, heat the water to 100.0°C, and then vaporize it to steam at 100.0°C. Remember break it into steps. The specific heat capacity of water is 4.184 J/g°C , the heat of vaporization of water is 40.6 kJ/18 g and the heat of fusion for ice is 6.02 kJ/18 g.

Step 1: Melt the ice

$$\frac{15 \text{ grams}}{} \left| \frac{6.02 \text{ kJ}}{18 \text{ grams}} \right. = 5.01667 \text{ kJ}$$

Step 2: Heat the liquid water from 0°C to 100°C

Q= S x M x ΔT

Q= (4.184)(15)(100)

Q= 6276 J

I want my answer in kilojoules, so I might as well convert this step's answer now:

$$\frac{6276 \text{ J}}{} \left| \frac{1 \text{ kJ}}{1\text{x}10^3 \text{ J}} \right. = 6.276 \text{ kJ}$$

Step 3: Vaporize the water to steam

$$\frac{15 \text{ grams}}{} \left| \frac{40.6 \text{ kJ}}{18 \text{ grams}} \right. = 33.83333 \text{ kJ}$$

Step 4. Add energy from each step to get total energy needed

5.01667 + 6.276 + 33.83333 = 45.126 kJ rounded to correct significant figures= 45 kJ