Chapter 3 – The Properties of Matter and Energy

Section 3-1: The Physical and Chemical Properties of Matter

Matter - occupies space and has mass
Mass - the amount of matter something contains
Substance – A particular kind of matter that has a uniform and definite composition.
All samples of a substance have identical physical properties.

The **physical properties** *of a substance are those that can be observed or measured without changing the substance into another substance.*

1) Sec 3-1.1 – The Physical States of Matter

There are three common states of matter: solid, liquid and gas.

Solids: A solid is rigid. It has its own shape and volume.
Liquids: A liquid flows. It has its own volume but will take the shape of its container.
Gases: A gas also flows. It will take both the shape and volume of its container.

Vapor describes a substance that, although in the gaseous state, is generally a liquid or solid at room temperature

Physical Property – A quality or condition of a substance that can be observed or measured without changing the substance's composition. Some examples of physical properties are:

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1. Color
2. Physical State (solid, liquid, or gas)
3. Odor (What is essential in order for something to have an odor?)
4. Taste (NEVER taste a chemical as a means to identification.
5. Freezing Point or Melting Point (same temperature) – temperature at which both the solid and liquid states are in equilibrium at normal atmospheric pressure.
6. Boiling Point or Condensation Point (same temperature) – temperature at which both the liquid and the gas states are in equilibrium at normal atmospheric pressure.
7. Ability to sublime (to change directly from a solid into a gas)
8. Density - the ratio of a substance's mass to its volume
9. Solubility

All matter ultimately consists of tiny particles that come together. There is a great diversity of matter. However, all matter is ultimately composed of a small number of fundamental particles (protons, neutrons and electrons).

2) Sec 3-1.2 – Changes in Physical State

Physical changes are changes in one or more physical properties but show no change in the fundamental substance. Such things as changes of state are physical changes because they do not affect the composition of the substance.

The **Freezing Point or Melting Point** (same temperature) – temperature at which both the solid and liquid states are in equilibrium at normal atmospheric pressure.

The **Boiling Point or Condensation Point** (same temperature) – temperature at which both the liquid and the gas states are in equilibrium at normal atmospheric pressure.

Pure substances will either melt or freeze, or boil and condense at exactly the same temperatures.

A **physical change** *in a substance does not involve a change in the composition of the substance but is simply a change in physical state or dimensions.*

3) Sec 3-1.3 – Types of Physical Properties

Intensive properties, such as clarity, color, density, etc., are those properties whose value does not depend upon the amount of material present.

Extensive properties, such as mass and volume, are those properties whose value does depend upon the amount of material present.

Both types of properties can be used to identify a substance, but *intensive* properties are more definitive.

4) Sec 3-1.4 – Chemical Changes and Chemical Properties

Substances have physical properties. One can also describe a substance in terms of its ***Chemical Properties*** which refer to the ability to form new substances.

A ***chemical change*** involves a change in the fundamental components of a substance. The substance changes into a different substance or substances. Chemical changes are often called ***chemical reactions***. When a chemical reaction occurs, one can often identify this by a change in physical properties: Is it now a solid when it was a liquid? Does it now smell? Has there been a color change?

5) Sec 3-1.5 – Chemical Change and Conservation of Mass

When a chemical change occurs the total mass of the elements and compounds involved does not change, only the identities of the substances involved. The Law of Conservation of Mass states that matter is neither created nor destroyed in a chemical reaction, ie., there is no change in mass during a chemical reaction.

Example – If 100 g of calcium carbonate, $CaCO_3$, when heated decomposes to produce 56 g of calcium oxide, CaO. How many grams of carbon dioxide, CO_2, escape into the air? $CaCO_3 \rightarrow CaO + CO_2$

Section 3-2: Density – A Physical Property

Earlier we described **density** as a derived measurement. It is also an intrinsic property of a substance. Density *is the ratio of the mass (usually in grams) to the volume (usually in milliliters for solids and liquids and liters for a gas).*

6) Sec 3-2.1 – Density as a Physical Property

It does not make a difference how much of a substance one has, the density will be the same for all quantities of that substance.

If something floats, its density is lower than that of the liquid. If it sinks, the density is greater.

7) Sec 3-2.2 – Density as a Conversion Factor

Density is not only a physical property. It can also be used to convert a mass of a substance to its volume or volume to its mass. *(This latter will be useful in later calculations.)*

8) Sec 3-2.3 – Specific Gravity

In certain circumstances, the term **specific gravity** is used to describe the density of a liquid. **Specific Gravity** is defined as the ratio of the density of a given liquid to the density of water at the same temperature or sometimes simply as water at 4°C. However, if the water is at 4°C then the specific gravity becomes the density of the substance because the density of water is defined as 1.0 at 4°C. Because it is a ratio of densities, specific gravity has no units.

Section 3-3: The Properties of Mixtures

Mixtures consist of a physical blend of two or more substances. They have variable composition. The parts of the mixture retain their characteristics and can be separated by physical means.

9) Sec 3-3.1 – Heterogeneous Mixtures

Heterogeneous mixture is not uniform in composition and contains two or more different substances.

10) Sec 3-3.2 – Homogeneous Mixtures and Solutions

A **Homogeneous mixture** has a completely uniform composition. Any small portion has the same composition as the whole. A **solution** is what chemists call a homogeneous mixture that is liquid.

11) Sec 3-3.3 – Alloys – Homogeneous Mixtures of Metals

Mixtures of metals are called **alloys**. Alloy usually exit in the solid phase. The composition of alloys is often expressed as a percent composition.

NOTE: *Pure substances* can either be elements or compounds. They always have the same composition.

There is a definite difference between *compounds* and *mixtures*. A compound only contains particles of identical composition and so is always the same. A compound can only be separated by chemical means, and it is separated into elements and/or simpler compounds. A mixture is a collection of compounds and/or elements that are present in varying amounts, and which can be separated by physical means.

12) Some techniques Used to Separate the Components of a Mixture

a) Filtration – Used to separate insoluble solids from liquids.
b) Distillation – Used to separate soluble solids from liquids, or liquids from liquids.

Section 3-4: The Forms and Types of Energy

Energy is the capacity or the ability to do work or to supply heat. The study of energy is called **Thermodynamics**.

13) Sec 3-4.1 – Forms of Energy

There are many forms or kinds of energy:

a. nuclear energy
b. light or radiant energy
c. chemical energy
d. heat energy
e. mechanical energy
f. electrical energy

Energy changes are subject to the same kind of law as matter changes in chemical reactions. The **Law of Conservation of Energy** states that energy cannot be created or destroyed but only transformed from one form to another. This law is most often called the **First Law of Thermodynamics**.

14) Sec 3-4.2 – Exothermic and Endothermic Changes

The **system** is the part of the universe that we are going to focus on. The **surroundings** include everything else in the universe.

In an **exothermic** reaction/process we feel/measure heat because the system is evolving heat and giving that heat to its surroundings. That is, heat flows out of the system. It feels **HOT**.

In an **endothermic** reaction/process the system is absorbing heat from the surroundings. That is, heat flows into the system. It feels **COLD**.

An important relationship to remember is that the energy gained (or lost) by the system must be equal to the energy lost (or gained) by the surroundings. This relationship will be important as we continue.

15) Some Additional Thoughts on Temperature and Heat

a) Temperature is the measure of the random motions of the components of a substance.
b) Heat is the flow of energy due to a temperature difference.

In talking about temperature: It does not make any difference whether we use Celsius or Kelvin for the temperature scale because what is important is the temperature **difference**. So, since the size of the degree for each is the same (remember Chapter 2), either scale can be used. In this case if the difference is 0 (no change in temperature), then no heat has been transferred

16) Sec 3-4.3 – Kinetic and Potential Energy

There are two types of energy:

a. **Kinetic Energy** is the energy produced from motion.
b. **Potential Energy** is the energy that is available because of position or composition.

Section 3-5: Energy Measurement and Specific Heat

Specific Heat (or specific heat capacity) is defined as the amount of heat required to raise the temperature of exactly 1 g of a substance exactly 1 °C (or K, kelvin). It is an intensive physical property.

17) Sec 3-5.1 – Units of Heat Energy

Units of heat energy are based on the specific heat of water. A **calorie** is the amount of heat required to raise the temperature of 1 gram of water from 14.5 to 15.5 °C. The unit of heat energy most often used in chemistry is the SI unit called the **joule**. The calorie is now defined in terms of joules.

1 cal = 4.184 joules (J) *(exactly)*

Now we have the specific heat of water in joules: 1.000 cal/g-°C = 4.184 J/g-°C.)

Since we are talking about a *temperature change* and not a specific measurement, it makes no difference whether Celcius or Kelvin scales are used. The change in temperature is represented as ΔT.

The formula to calculate specific heat is usually given as

$$\text{Specific heat} = \frac{\text{amount of heat energy (J or cal)}}{\text{mass (g)} \times \Delta T\ (^\circ\text{C})}$$

18) Sec 3-5.2 – The Nutritional Calorie and Heat Exchange

The energy we get from food can be measured and expressed in terms of calories. The nutritional calorie is actually one *kilocalorie* (10^3 cal) as we discussed above. It is often referred to as a Calorie (1 Cal) – or "big C" calorie.

The principle of heat exchange says that heat lost equals heat gained as long as no heat is lost to the surroundings. Generally we describe this as:

Heat lost by A = Heat gained by B

$$q_{out} = -q_{in}$$

Some people do not use the minus sign when they talk about this relationship, they merely remember that on one side heat is going out and on the other heat is going in, and that the amount of heat is really the same.

And we can then show that:

Energy (heat) (q) = Specific heat (s) x Mass (grams) (m) x Temperature change (ΔT)

or

$$q = s * m * \Delta T$$

19) Calorimetry

Calorimeters are devices used to measure the amount of heat absorbed or released during chemical or physical processes. Calorimetry is the accurate and precise measurement of the heat change for chemical and physical processes.

Examples

a) A chemistry student dissolves 4.51 grams of sodium hydroxide in 100.0 mL of water at 19.5°C (in a calorimeter cup). As the sodium hydroxide dissolves, the temperature of the surrounding water increases to 31.7°C. Determine the heat of solution of the sodium hydroxide in J/g.

b) Calculate the final temperature when 50.0 g of Al (s = 0.899 J/g-°C) which is at 100°C is placed in 80 g of water which is at 20°C.

c) A student has 37.0 g of water at a given temperature and 35.09 g of a metal at a higher temperature. If the ΔT of the water is 1.8°C and the ΔT of the metal is 77.1°C, what is the specific heat of the metal?