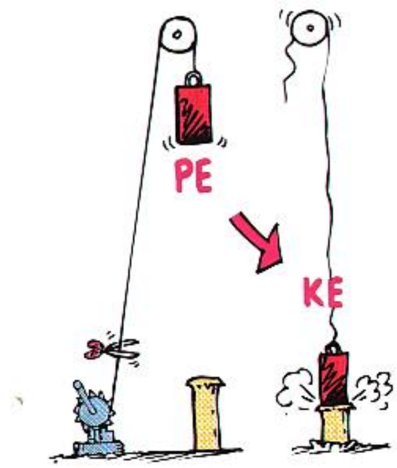
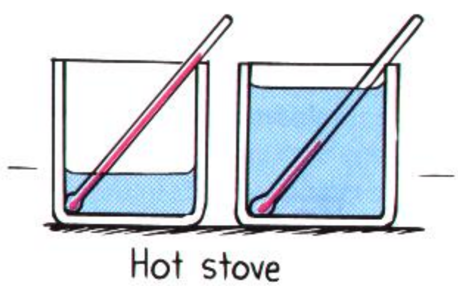
ENERGY & HEAT



- ***Energy*** is defined as the capacity of matter to do ***work***.
- There are two types of energy:
  1. ***Potential*** (stored)
  2. ***Kinetic*** (moving)
- Energy possesses many ***forms*** (chemical, electrical, thermal, etc.), and can be ***converted*** from one form into another.
- In chemistry, ***energy*** is commonly expressed as ***heat***.
- ***Heat*** is measured in SI units of ***joule*** or the common unit of ***calorie***. (1 cal=4.184 J)

<u>Heat vs. Temperature:</u>

- Although the ***same*** amount of ***heat*** is added to both containers, the ***temperature increases*** more in the container with the ***smaller amount*** of water.



- Heat is a form of energy associated with particles of matter. Heat is the total energy of all particles of matter.
- Temperature is a measure of the intensity of heat or how hot or cold a substance is. Temperature is the average kinetic energy of particles of matter.

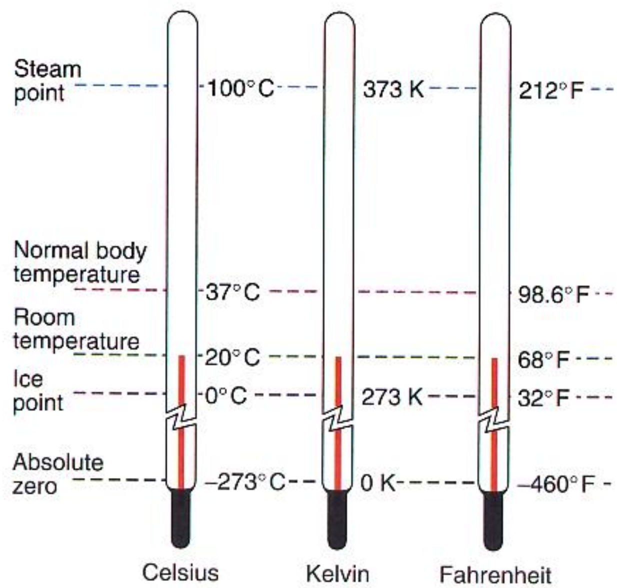
TEMPERATURE SCALES

- ***Temperature*** is the measure of ***how hot or cold*** a substance is.

- ***Thermometer*** is an instrument that measures temperature and is based on ***thermometric properties*** (i.e. expansion of solids or liquids, color change, etc.) of matter.

- Three ***scales*** are used for measuring temperature:

    1. ***Fahrenheit*** (32 - 212)
    2. ***Celsius*** (0 - 100)
    3. ***Kelvin (absolute)*** (273 -373)



- To convert from one scale to another, the following relationships can be used:

    K = °C + 273
    °F = (1.8 x °C) + 32
    °C = (°F–32)÷1.8

or alternately,

    °F = [(°C+40) x 1.8]–40
    °C = [°F+40) ÷ 1.8]–40

<u>Examples:</u>

1. The melting point of silver is 960.8°C. Convert to Kelvin.

    K = °C + 273 K =

2. Pure iron melts at about 1800 K. What is this temperature in °C?

    °C = K – 273 C =

3. On a winter day the temperature is 5°F outside. What is this temperature on the Celsius scale?

    °C = [°F+40) ÷ 1.8]–40 =

4. To make ice cream, rock salt is added to crushed ice to reach a temperature of –11°C. What is this temperature in Fahrenheit?

SPECIFIC HEAT

- Different materials have different ***capacities*** for storing heat.

- The ***specific heat*** of a substance is the ***amount of heat*** required to change the temperature of ***1 g*** of that substance by ***1 °C***.

- Units of specific heat are **J/g°C** or **cal/g°C**.

- Shown below are the specific heat of some substances:

The filling of hot apple pie may be too hot to eat, whereas the crust is not.

Substance	**(cal/g°C)**	**(J/g°C)**
Aluminum	0.214	0.897
Copper	0.0920	0.385
Iron	0.0308	0.129
Ammonia	0.488	2.04
Ethanol	0.588	2.46
Water	1.00	4.184

- When heated, substances with ***low specific heat*** get hot ***faster*** while substances with ***high specific*** heat get hot at a ***slower*** rate.

- When cooled, substances with ***low specific heat*** get cool ***faster*** while substances with ***high specific*** heat cool at a ***slower*** rate.

- The amount of heat lost or gained by a system is determined by the following equation:

$$\textbf{Heat} = \begin{pmatrix}\textbf{mass of}\\ \textbf{substance}\end{pmatrix} \begin{pmatrix}\textbf{specific heat}\\ \textbf{of substance}\end{pmatrix} \begin{pmatrix}\textbf{change in}\\ \textbf{temperature}\end{pmatrix}$$

$$\downarrow \qquad\qquad \downarrow \qquad\qquad \downarrow$$

$$\mathbf{Q = (m) \times (s) \times (\Delta T)}$$

Examples:

1. Determine the amount of heat needed to raise the temperature of 200. g of water by 10.0 °C. (Specific heat of water is 4.184 J/g°C)

   m =

   s =

   ΔT =

   Q =

2. Calculate the specific heat of a solid if 1638 J of heat raises the temperature of 125 g of the solid from 25.0 to 52.6 °C.

   m =

   s =

   ΔT =

   Q =

3. Ethanol has a specific heat of 2.46 J/g°C. When 655 J are added to a sample of ethanol, its temperature rises from 18.2°C to 32.8°C. What is the mass in grams of the ethanol sample?

   m =

   s =

   ΔT =

   Q =

ENERGY & NUTRITION

- The foods we eat provide energy for our bodies. Vitamins and minerals are necessary for health but have little energy value.
- Carbohydrates are the main source of fuel for the body, but when their reserves are exhausted, fats and then proteins can be used for energy.
- In the field of nutrition, the energy from food is measured in units of ***Calories (Cal)***. One Calorie is equal to ***1000 calories or 1 kilocalorie (kcal)***.
- In the laboratory, foods are burned in a calorimeter to determine their energy. A sample of food is burned in the calorimeter, and the energy released is absorbed by water surrounding the calorimeter. The energy of the food can be calculated from the mass of the food and the temperature increase of the water.

<u>Examples:</u>

1. A 2-oz serving of pasta provides 200 Cal. What is the energy value of pasta in Cal/g?

Step 1: Given 200 Cal/2 oz Need Cal/g

Step 2: oz $\xrightarrow{\text{English – English factor}}$ lb $\xrightarrow{\text{Metric – English factor}}$ g

Step 3: $\dfrac{16\text{ oz}}{1\text{ lb}}$ and $\dfrac{454\text{ g}}{1\text{ lb}}$

Step 4:

$$\frac{200\text{ Cal}}{2\text{ oz}} \times \frac{16\text{ oz}}{1\text{ lb}} \times \frac{1\text{ lb}}{454\text{ g}} = 4\text{ Cal/g}$$

2. A 2.3-g sample of butter is placed in a calorimeter containing 1900 g of water at a temperature of 17°C. After the complete combustion of the butter, the water has a temperature of 28°C. What is the energy value of butter in Cal/g?

m =

s =

ΔT =

Q =

ENERGY CONTENT OF FOOD

- The energy value of foods are the kilocalories or kilojoules obtained from the complete combustion of 1 g of a carbohydrate, fat or protein.

Food Type	**kJ/g**	**Cal/g (kcal/g)**
Carbohydrate	17	4
Fat	38	9
Protein	17	4

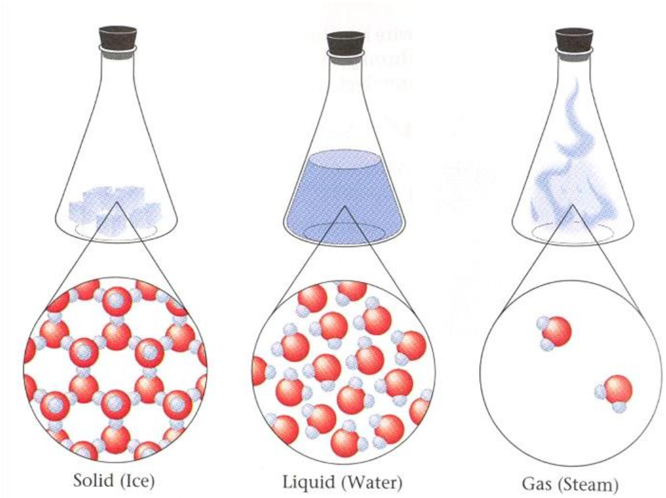
Example:

1. What is the energy value (in Cal) for a piece of chocolate cake that contains 34 g of carbohydrate, 10 g of fat and 5 g of protein? (Calculate answer to 2 sig figs.)

2. A 1-oz (28 g) serving of oat-bran hot cereal with half a cup of whole milk contains 22 g of carbohydrate, 7 g of fat, and 10 g of protein. If you eat two servings of the oat bran for breakfast, how many kilocalories will you obtain? (Calculate answer to 2 sig figs)

CLASSIFICATION OF MATTER

- ***Matter*** is anything that has ***mass***, and occupies ***space***.
- Matter can be classified by its *physical state* as **solid**, **liquid** or ***gas***.



Solid:
- ***Densely*** packed matter with ***definite shape*** and ***volume***.
- Particles have ***strong forces*** of attraction towards each other.
- Solids are **not very compressible**

Liquid:
- ***Loosely*** packed matter with ***definite volume*** but ***indefinite shape***.
- Particles have ***moderate forces*** of attraction towards each other and are **mobile**.
- Liquids are **slightly compressible**.

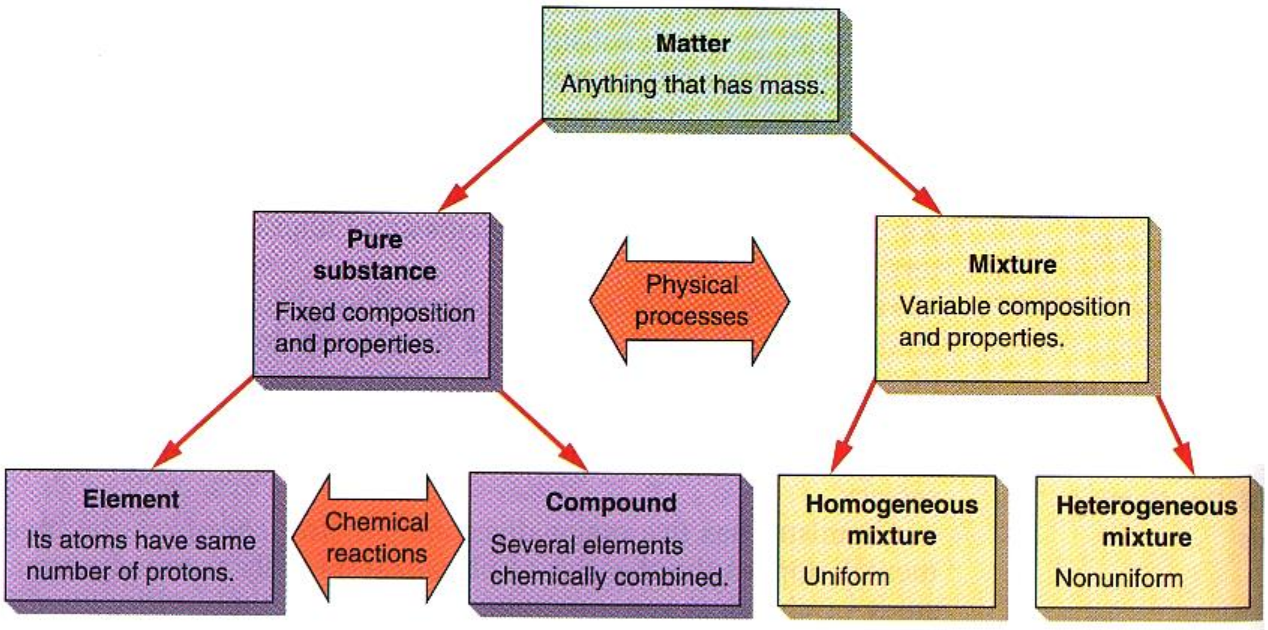
Gas:
- ***Very loosely*** packed matter with ***no definite shape*** or ***volume***.
- Particles have ***little or no forces of attraction*** towards each other.
- Gases are **very compressible**.

SUMMARY OF PROPERTIES OF MATTER

State	Shape	Volume	Particles	Compressibility
Solid	Definite	Definite	Densely packed	Very slight
Liquid	Indefinite	Definite	Mobile	Slight
Gas	Indefinite	Indefinite	Far apart	High

CLASSIFICATION OF MATTER

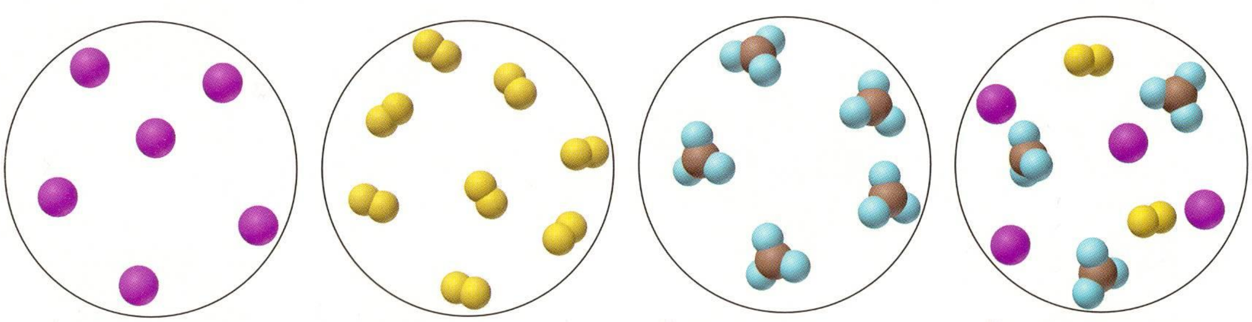
- Matter can also be classified by its ***composition*** as ***pure substance*** or ***mixture***.



Element:
- ***Pure substance*** that is made up of only ***one type of atom***.
- Examples include: gold, copper, hydrogen.

Compound:
- ***Pure substance*** that is made up of ***two or more elements chemically*** combined together.
- ***Properties are unique*** compared to its components.
- Smallest particle is a ***molecule***.
- Examples include: water, salt, aspirin.

Classify each of the following substances as element, compound or mixture.

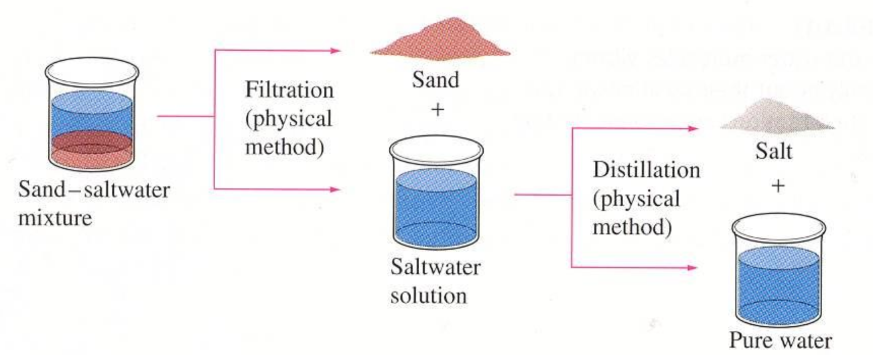


MIXTURES

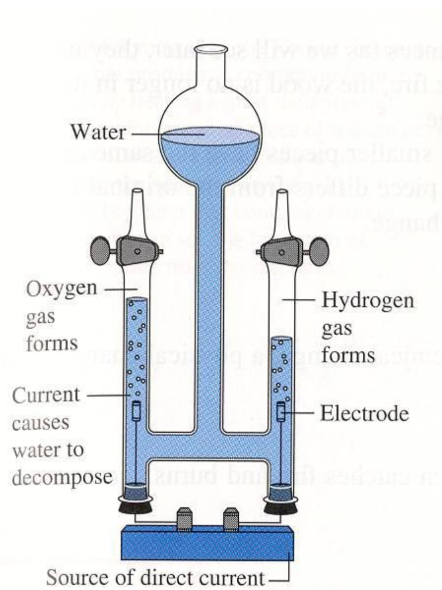
Mixture:
- ***Two or more substances physically*** combined together.
- ***Properties are similar*** to those of its components.
- Can be ***separated easily*** by a ***physical process.***
- Two types***: heterogeneous and homogeneous.***

Heterogeneous:
- Mixture that is ***non-uniform*** in composition.
- Examples include: vegetable soup, cement, salad dressing.

Homogeneous:
- Mixture that is ***uniform*** in composition**.**
- Commonly referred to as ***solution***.
- Examples include: gasoline, soda pop, salt solution.



Separation of a mixture through physical methods



Separation of a compound through chemical methods

PHYSICAL & CHEMICAL PROPERTIES

- The ***characteristics*** of a substance are called its ***properties***.

- ***Physical properties*** are those that describe the matter ***without changing its composition***. Examples are density, color, melting and boiling points, and electrical conductivity.

- ***Chemical properties*** are those that describe how matter behaves in combination with other matter, and involve ***change in its composition***. Examples are flammability, corrosion, and reactivity with acids.

Examples:

Identify each of the following properties as physical or chemical:

1. Oxygen is a gas

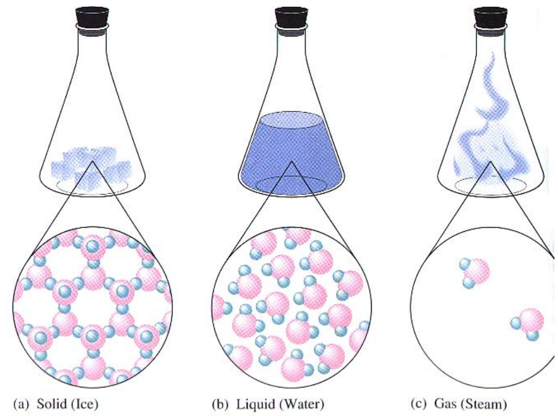
2. Helium is un-reactive

3. Water has high specific heat

4. Gasoline is flammable

5. Sodium is soft & shiny

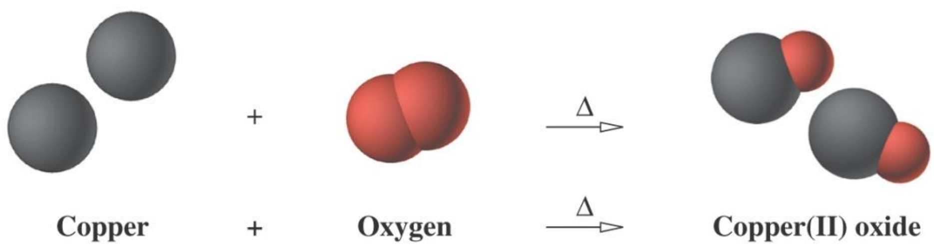
PHYSICAL & CHEMICAL CHANGES



- ***Changes in physical properties*** of matter that do not involve change in its composition are ***called physical changes***.

- Examples are melting, evaporation and other phase changes. Physical changes are ***easily reversible***.

- A change that ***alters the chemical composition*** of matter, and ***forms new substance*** is called a ***chemical change***.



- Examples are burning, rusting, and reaction with acids.

- Chemical changes are ***not easily reversible***, and are commonly called ***chemical reactions***.

<u>Examples:</u>

Identify each of the following changes as physical or chemical:

1. Cooking food

2. Mixing sugar in tea

3. Carving wood

4. Burning gas

5. Food molding

CHANGE OF STATE

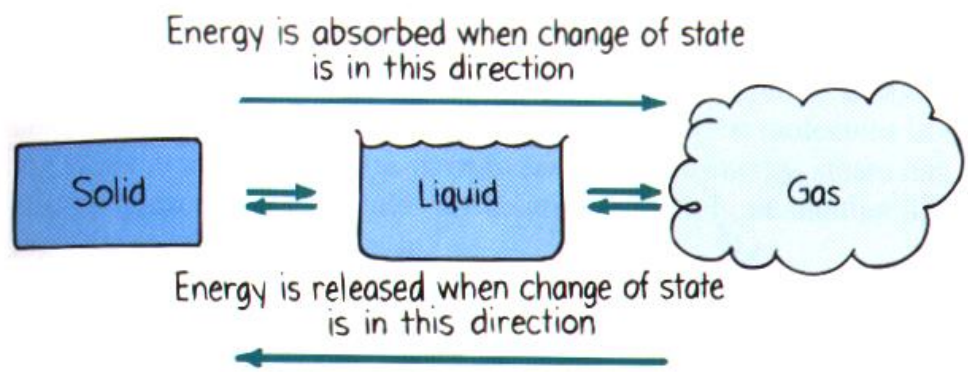
- When matter ***releases*** or ***absorbs energy without a change in temperature***, ***phase change*** occurs (e.g. melting, evaporation).

- The common phase changes are as follows:

$$\text{solid} \underset{\text{freezing}}{\overset{\substack{\text{fusion}\\\text{melting}}}{\rightleftharpoons}} \text{liquid}$$

$$\text{liquid} \underset{\text{condensation}}{\overset{\text{evaporation}}{\rightleftharpoons}} \text{gas}$$

$$\text{solid} \underset{\text{deposition}}{\overset{\text{sublimation}}{\rightleftharpoons}} \text{gas}$$

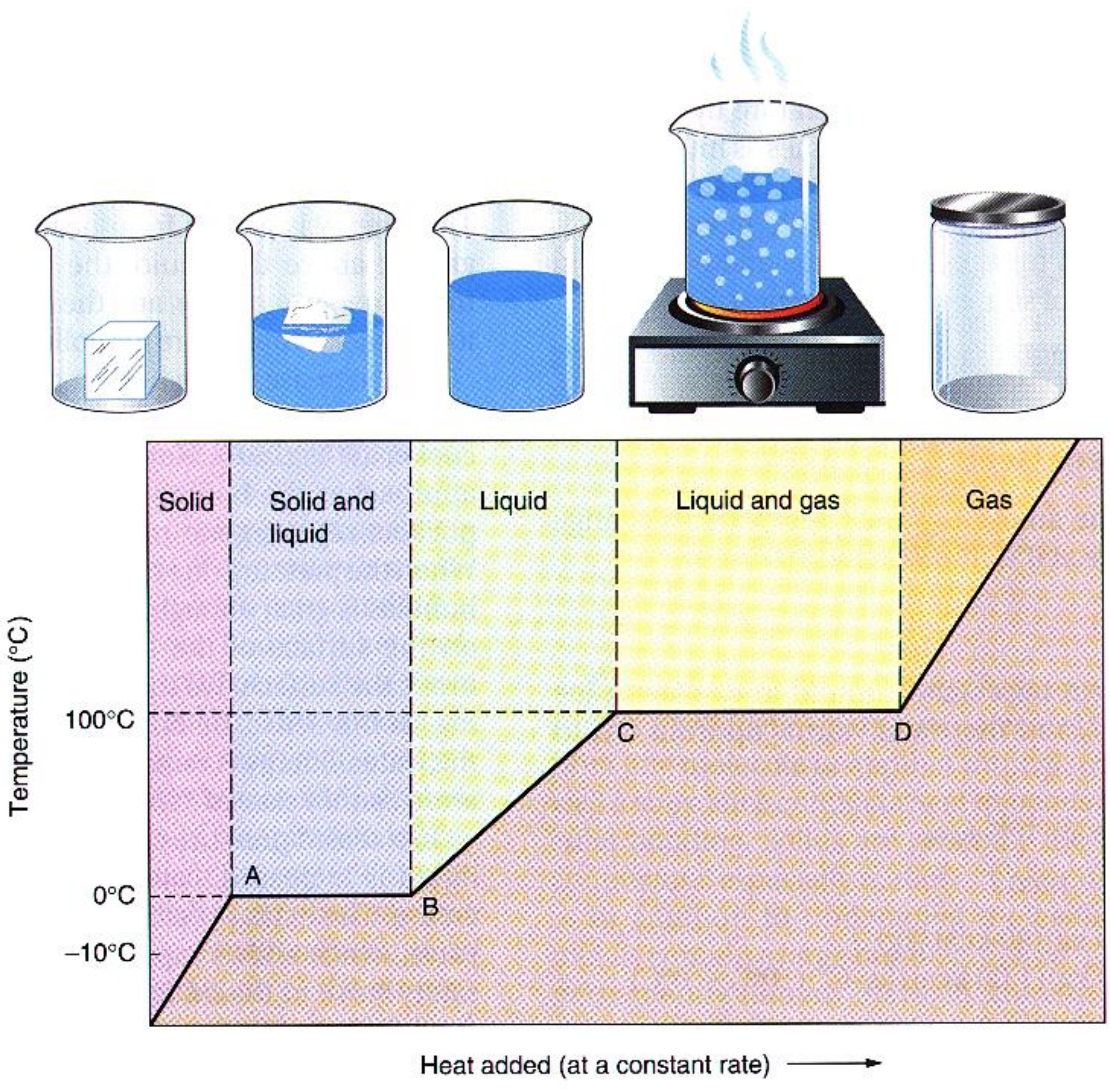


- Phase changes that involve **absorption of heat are cooling processes**.

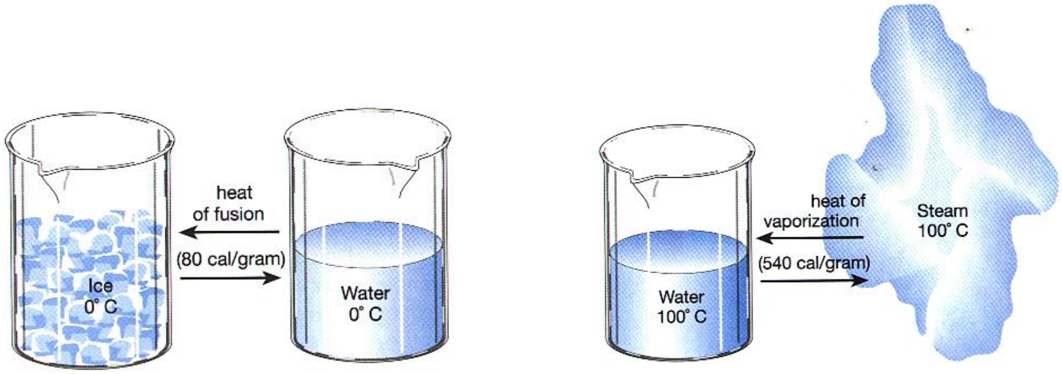
- Phase changes that involve **release of heat are warming processes**.

HEAT & COOLING CURVES

- When ***heat is added*** to ice, it ***absorbs*** the heat ***without a change in temperature***, causing a ***phase change***.
- Similarly, when ***heat is added*** to hot water, a ***phase change*** occurs ***without an increase in temperature***.

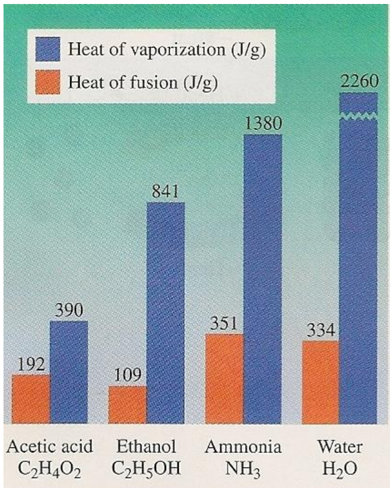


HEAT OF FUSION & VAPORIZATION



Heat of Fusion (H_f)	Heat of Vaporization (H_v)
The quantity of heat required to melt 1 g of solid	The quantity of heat required to evaporate 1 g of liquid

- For any substance the heat of vaporization is greater than the heat of fusion.



- The amount of ***heat***, released or absorbed during phase change, depends on the ***amount*** of substance and the ***heat of vaporization*** or ***heat of fusion***.

Q = mass x heat of fusion

Q = mass x heat of vaporization

Examples:

1. How much heat is required to melt 50 g of ice at 0°C? Heat of fusion for ice is 80 cal/g.

   m =

   H_f =

   Q = ???

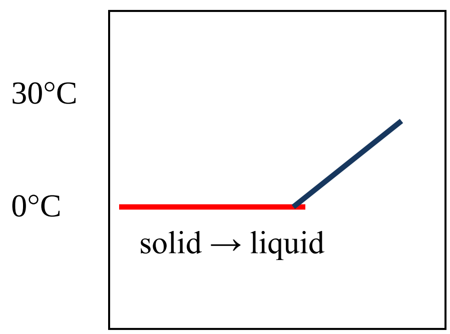
2. How much heat is required to vaporize 50.0 g of water at 100°C? Heat of vaporization for water is 540 cal/g.

   m =

   H_v =

   Q = ???

3. Calculate the amount of heat required to change 25 g of ice at 0°C to water at 30.0°C.



$Q_{total} = Q_{melt\ ice} + Q_{change\ T}$

$Q_{melt\ ice} = m \times H_f =$

$Q_{change\ T} = m \times s \times \Delta T =$

$Q_{total} =$