Chapter 1:

Matter and Energy

Chapter 1 Topics:

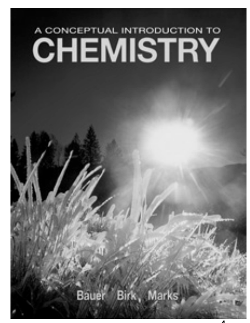
1. Matter and its Classification
2. Physical and Chemical Changes and Properties of Matter
3. Energy and Energy Changes
4. Scientific Inquiry

Chapter 1 Math Toolboxes:

1.1 Scientific Notation
-- Also called exponential notation
1.2 Significant Figures
1.3 Units and Conversions

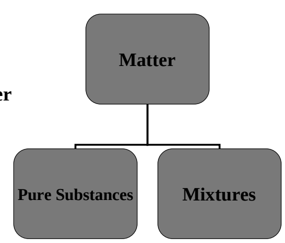
1.1 Matter and it's Classification

- **Matter is anything that occupies space and has mass.**
- **Forms of energy are NOT matter. Heat and light, for example, do not occupy space and have no mass.**
- **Consider the interaction between matter and energy in this picture.**



Composition of Matter

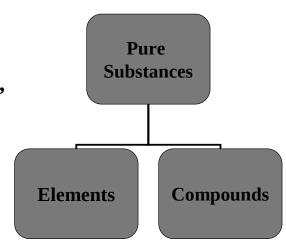
- **We classify matter so that we can understand it better.**
  - **One way to classify matter is as pure substances or mixtures.**



Composition of Matter

Pure Substances:

- **have the same composition throughout, and from sample to sample.**
- **can be further classified as either elements or compounds.**

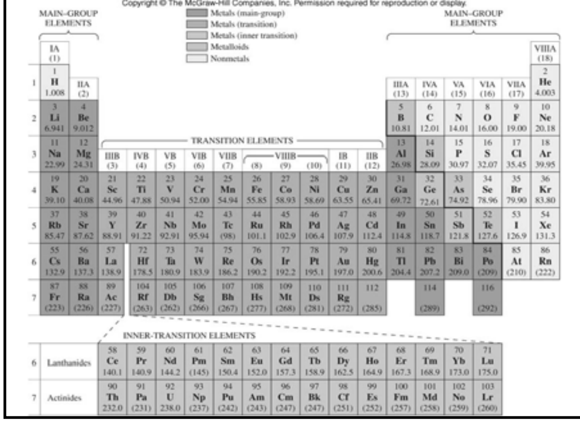


Elements

- **An element is a substance that cannot be broken down into simpler substances even by a chemical reaction.**
- **All known elements are organized on the periodic table.**

7

Figure 1.03



Elements and their Symbols

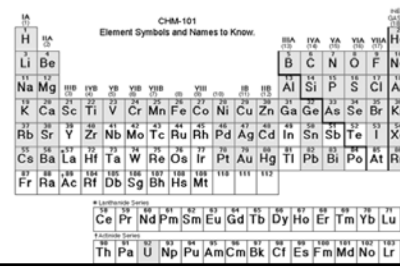
- Element symbols often consist of one or two letters of the element's name.
- Examples: carbon: C calcium: Ca
- How do we explain that Fe is the symbol for iron?

TABLE 1.1 Symbols of Selected Elements

English Name	Original Name	Symbol	English Name	Original Name	Symbol
Copper	Cuprum	Cu	Potassium	Kalium	K
Gold	Aurum	Au	Silver	Argentum	Ag
Iron	Ferrum	Fe	Sodium	Natrium	Na
Lead	Plumbum	Pb	Tin	Stannum	Sn
Mercury	Hydragyrum	Hg	Tungsten	Wolfram	W

Elements and their Symbols

- **Know the names and symbols of the most common elements indicated in the handout.**

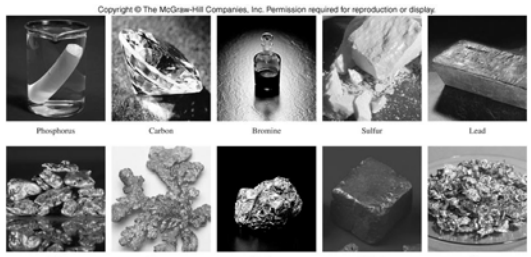


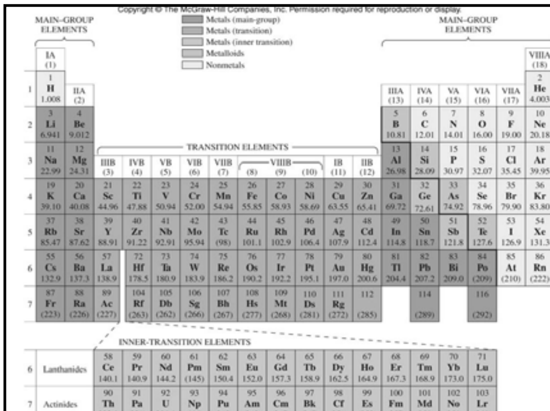
pertabinfo.swf

10

Some Elements

- Which are metals and which are nonmetals?





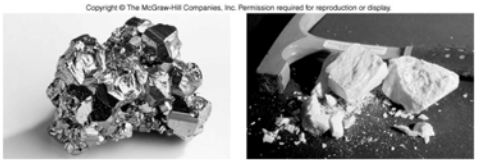
12

Figure 1.03

Compounds

- A compound is a pure substance composed of two or more elements combined chemically in definite proportions.
- A compound has properties that are different from those of its component elements.

Fig. 1.5



Compound: Iron pyrite, Fe_2S_3 **Elements: Iron, Fe, and Sulfur, S**

Compound

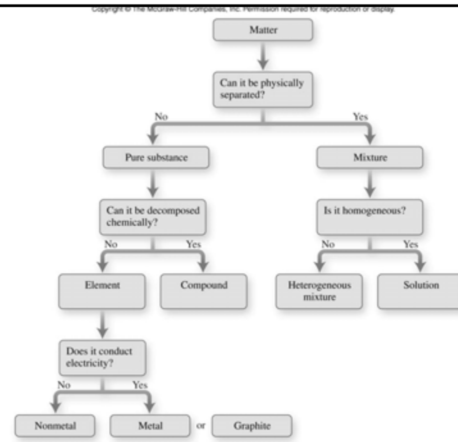
- Pure sand is the compound silicon dioxide, SiO_2.

Figure 1.01

Elements and Compounds

- Identify each of the following as an element or compound.

1. He
2. H_2O
3. sodium chloride
4. copper



classify-matter.swf

Mixtures

- A mixture is a combination of two or more elements or compounds .
- Mixtures differ from pure compounds in that that their components can be separated by physical processes.

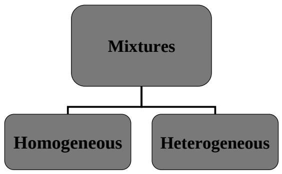
- Examples:
  - Pencil lead
  - Salt water
  - Air

Salt Being Separated by Evaporation from the Great Salt Lake

Figure 1.06

Mixtures

- **Mixtures can be further classified as homogeneous and heterogeneous.**
- **Homogeneous mixtures have the same composition throughout.**
- **Heterogeneous mixtures do not.**



19

Mixtures

- Classify each of the following mixtures as **homogeneous** or **heterogeneous**:
  - Salt water
  - Lake water
  - Tap water
  - Air
  - Brass (an alloy of copper and zinc)
  - Potting soil
  - Cake mix

20

Representations of Matter

- **Macroscopic** – we can see with our eyes
- **Molecular level** – a magnification to a level that shows atoms
- **Atom** – the smallest unit of an element; represented as single sphere.
- **Molecule** – two or more bound atoms

21

Molecular-Level Representations of Matter – Copper Atoms

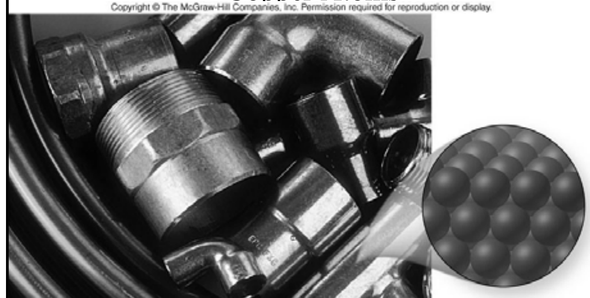


Figure 1.08

22

Molecular-Level Representations of Matter – Helium Atoms

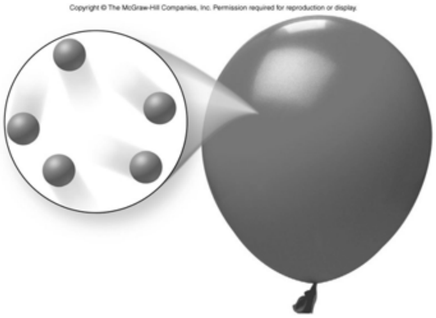


Figure 1.09

23

Molecular-Level Representations of Matter – water molecules

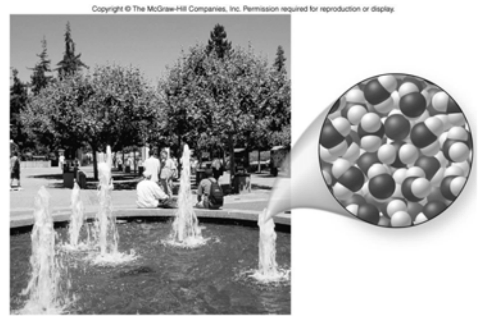


Figure 1.10

24

Molecular-Level Representations

- Does this image represent atoms or molecules?
- Is this an element, compound, or mixture?

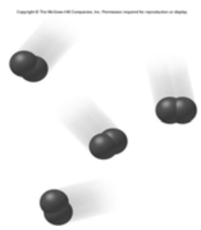


Figure 1.11

25

Different Ways to Represent Water

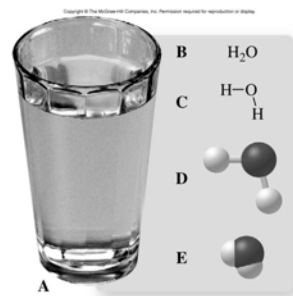
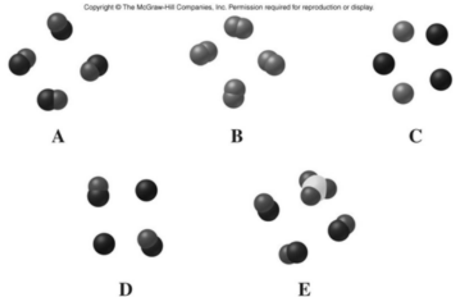


Figure 1.12

26

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



27

States of Matter

- A different way to classify matter is by its physical state: solid, liquid, or gas.
- What are the macroscopic properties of each?
- How do the atoms and molecules of solids, liquids, and gases behave differently?

Change of State

28

Solid and Liquids States of Copper

Figure 1.14

29

Gases can be Compressed

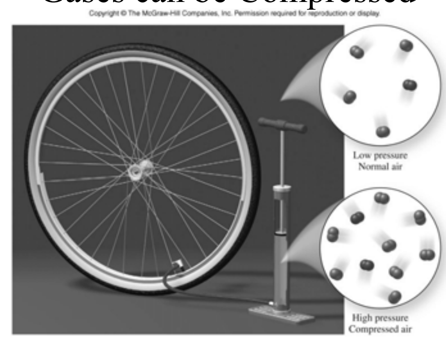


Figure 1.15

30

Water vapor condenses from the air onto the cold surface of the glass.

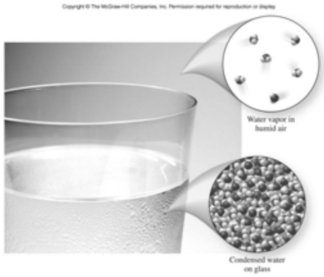


Figure1.16

31

TABLE 1.3 Symbols for Physical State

Physical State	Symbol	Example (bromine)
Solid	(*s*)	$Br_2(s)$
Liquid	(*l*)	$Br_2(l)$
Gas	(*g*)	$Br_2(g)$
Aqueous (dissolved in water)	(*aq*)	$Br_2(aq)$

Table 1.03

32

1.2 Physical and Chemical Changes and Properties of Matter

- A physical property is a characteristic that we can observe without changing the composition of a substance.
- Examples
  - Color
  - Odor

  qualitative properties

  - Mass
  - Volume
  - Density
  - Temperature

  quantitative properties

33

Math Toolbox 1.3

- **Units and Conversions**

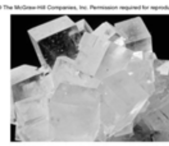
<u>Metric Base Units and Derived Units</u>

- Length: meter (m)
- Mass: kilogram (kg)
- Time: second (s)
- Temperature: Kelvins (K)
- Mole: mol

34

Mass

- We usually measure the mass of an object by weighing it on a balance.
- How many milligrams in a gram?
- How many grams in a kilogram?



Mass: 50 mg, 0.05 g, or 5×10^{-5} kg

Mass: 7×10^{7} mg, 7×10^{4} g, or 70 kg

35

Conversion (Math Toolbox 1.3) (& back cover of text)

<u>Prefixes (Table 2.3)</u>			<u>Length measurements</u>
giga-	G	10^9	1 Gg = 10^9 g
mega-	M	10^6	1 Mg = 10^6 g
kilo-	k	10^3	1 kg = 10^3 g
centi-	c	10^{-2}	1 cg = 10^{-2} g
milli-	m	10^{-3}	1 mg = 10^{-3} g
micro-	μ	10^{-6}	1 μg = 10^{-6} g
nano-	n	10^{-9}	1 ng = 10^{-9} g
pico-	p	10^{-12}	1 pg = 10^{-12} g

36

Conversion
(Math Toolbox 1.3)
(& back cover of text)

Prefixes (Table 2.3) Length measurements

giga-	G	10^9	1 Gm = 10^9 m
mega-	M	10^6	1 Mm = 10^6 m
kilo-	k	10^3	1 km = 10^3 m
centi-	c	10^{-2}	1 cm = 10^{-2} m
milli-	m	10^{-3}	1 mm = 10^{-3} m
micro-	μ	10^{-6}	1 μm = 10^{-6} m
nano-	n	10^{-9}	1 nm = 10^{-9} m
pico-	p	10^{-12}	1 pm = 10^{-12} m

Mass Unit Conversions
(Math Toolbox 1.3)

- Convert 12.0 grams to milligrams.
- Convert 12.0 grams to ounces (1 oz = 28.34 g)

Math Toolbox 1.1

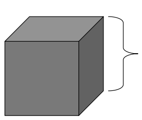
- **Scientific Notation**
  - Powers of Ten (Slide Show).
  - 0.000523 = 5.23×10^{-4}

Math Toolbox 1.2

- **Significant Figures**
  - All non-zero digits are significant. (435 g)
  - A zero that falls between two significant digits is significant. (405 g; 40.5 g)
  - Zeros to the right of a sig. digit and to the right of a decimal pt. are significant. (5.00 g)
  - Zeros to the left of the first significant digit are not significant. (0.151 g; 0.00405 g)
  - If a number is >1, the zeros to the right of the last nonzero digit may or may not be significant. Use scientific notation to specify.

Volume

- We can measure the volume of a cube by measuring the length of one of its sides, and then cubing the length. If the length of a side is 2.0 cm, what is the volume of this cube?



2.0 centimeters

Volume

- Volumes of liquids are usually measured in units of milliliters (mL).
- 1 mL = 1 cm^3 exactly
- How many mL in 1 L?

Some 250-mL, 500-mL, and 1-L containers

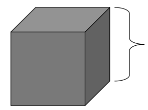
Volume Unit Conversions

- Convert 25.0 mL to L.
- Convert 25.0 mL to quarts (1 L = 1.057 qt)

43

Density

- The density of a substance is the ratio of its mass to volume:
- If the mass of the cube is 11.2 grams, what is its density?



2.0 centimeters

44

Density

- Which liquid is the most dense. Which is least dense?
- Compare the density of the water and the professor.

Figure 1.19

45

Gold has a greater mass than aluminum. Which cube has the greater density?

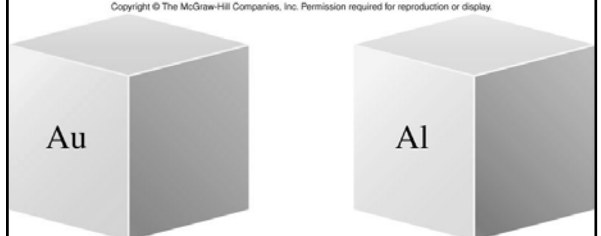
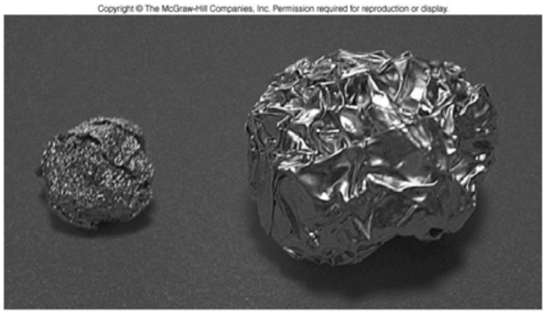


Figure 1.20

46

Given that these samples of metals have the same mass, which has the greater density?

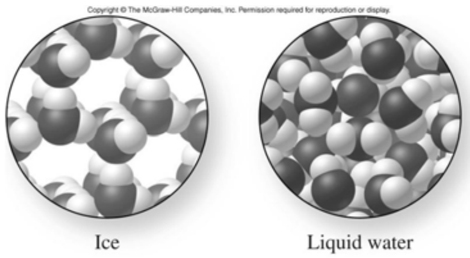


Density = mass/volume

- Why is regular soda more dense?

48

Why is ice less dense than liquid water?



Temperature

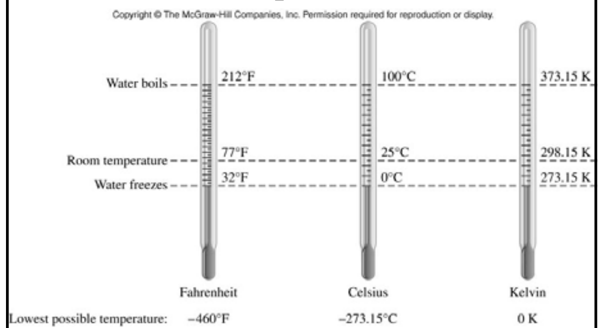


Figure 1.22

Temperature Scales

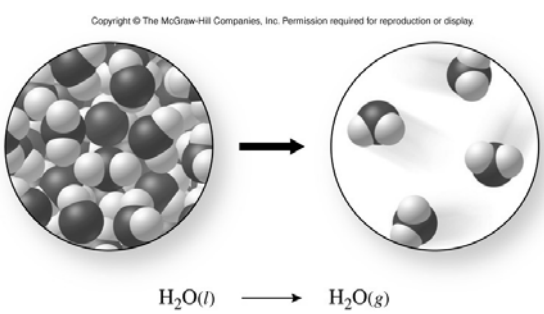
- $T_K = T_{°C} + 273.15$
- $T_{°F} = 1.8(T_{°C}) + 32$

- Boiling Point of Water:
  - 212°F, 100°C, 373.15 K
- Freezing Point of Water:
  - 32°F, 0°C, 273.15 K
- Lowest Possible Temperature:
  - −273.15°C, 0.00 K

Physical Changes

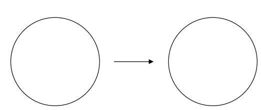
- A physical change is a process that changes the physical properties of a substance without changing its chemical composition.
- Phase changes are physical changes.

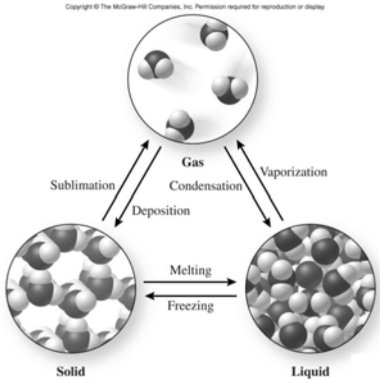
Vaporization or Evaporation



Sublimation of Dry Ice (CO_2)

- $CO_2(s) \rightarrow CO_2(g)$
- Draw a molecular-level representation for the sublimation of CO_2.





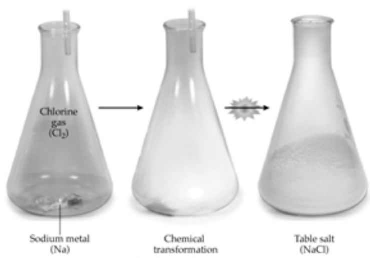
55

Chemical Changes

- A chemical change is a process where one or more substances are converted into one or more new substances. (Also called a chemical reaction)
- Examples:
  - Pennies tarnishing
  - Burning gasoline
  - The reaction of hydrogen and oxygen to form water

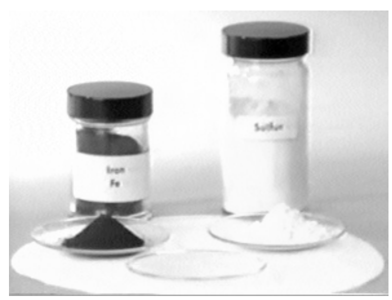
56

Chemical Changes



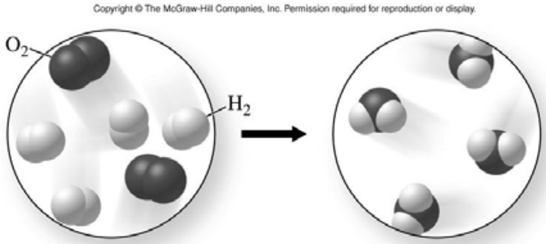
57

Chemical vs. Physical Changes



58

Chemical Change



$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$$

Figure 1.25

59

Chemical Properties

- **Chemical Properties are descriptions of the ability of a substance to undergo a chemical change.**
- **Examples:**
  - **Hydrogen burns easily with oxygen**
  - **Helium is unreactive**
  - **Iron rusts**
  - **Silver tarnishes**
  - **Gold is very unreactive**

60

Is Boiling Water a Chemical or Physical Change?

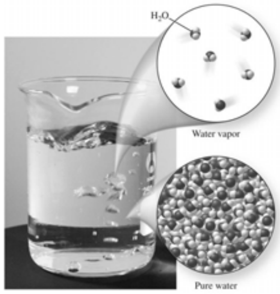
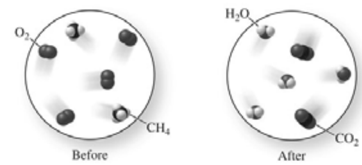


Figure 1.27b

EXAMPLE 1.11 **Physical and Chemical Changes**

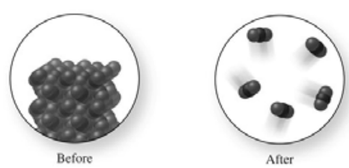
Do the following molecular-level images represent a chemical change or a physical change?



EXAMPLE 1.11 **Physical and Chemical Changes** *continued*

Practice Problem 1.11

Do the following molecular-level images represent a chemical change or a physical change?



Further Practice: 1.73 and 1.74

1.3 Energy and Energy Changes

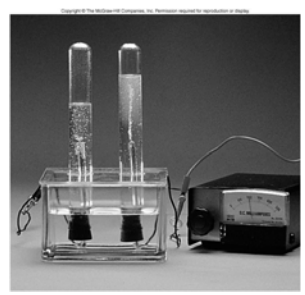
- **When chemical or physical changes occur, energy changes also occur.**
- **Some processes release energy and some require an energy input.**
- **Examples:**
  - **When wood burns with oxygen, energy in the form of heat is released.**
  - **When ammonium nitrate dissolves in water in a cold pack, energy in the form of heat is absorbed.**

When hydrogen burns with oxygen, energy in the form of heat is released.

rxn

Figure 1.28

Electricity is used to decompose water into its elements

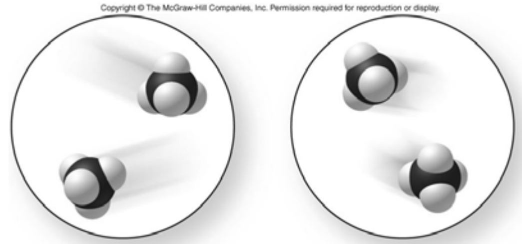


Energy

- **Kinetic energy – energy of motion**
  - **The kinetic energy of a sample will increase as temperature is increased.**
- **Potential energy – energy possessed by an object because of its position; stored energy**
  - **As a ball is raised up in the air, its potential energy increases.**
  - **Very reactive substances have high potential energy.**

67

Which pair of molecules has more kinetic energy?



68

Units of Energy

- **Calories (Cal); calories (cal); joules (J)**

- **The unit Calorie (Cal) is used to describe the energy content of food.**

1 Cal = 1000 cal = 1 kcal

4.184 J = 1 cal

69

➢ **One serving of this cereal has 190 Calories.**

➢ **What is the energy content in units of joules?**