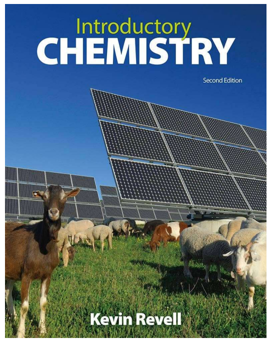
Introductory Chemistry
Chem 103

Chapter 7 – Mass Stoichiometry

Lecture Slides



Formula Mass and Percent Composition

1	H	1.01
6	C	12.01
8	O	16.00

formula mass the mass of a single molecule or formula unit.

"formula weight"
"molecular mass"
"molecular weight"

Ex.: What is the formula mass of a water molecule?

H_2O

$= 2(1.01\ u) + 1(16.00\ u)$

$= 18.02\ u$

Determining Formula Mass Practice

19	K	39.10
6	C	12.01
8	O	16.00

Potassium carbonate, K_2CO_3, is a common water-softening agent. What is the formula mass of this compound?

K_2CO_3

$= 2(39.10\ u) + 1(12.01\ u) + 3(16.00\ u)$

$= 138.21\ u$

Percent Composition

$$\text{percent composition of one element} = \frac{\text{mass of one element}}{\text{mass of compound}} \times 100\%$$

Determining Percent Composition Practice

Octane, a component of gasoline, has the molecular formula C_8H_{18}. What is the percent composition of carbon and hydrogen in octane?

mass C = 8(12.01 u) = 96.08 u

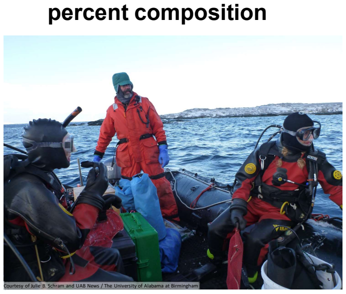
mass H = 18(1.01 u) = 18.18 u

mass of C_8H_{18} = 8(12.01 u) + 18(1.01 u) = 114.26 u

$$\% \text{ carbon} = \frac{\text{mass carbon}}{\text{total formula mass}} \times 100\% = \frac{96.08\ u}{114.26\ u} \times 100\% = 84.09\%$$

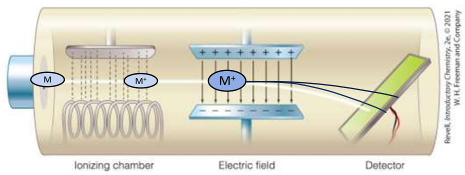
$$\% \text{ hydrogen} = \frac{\text{mass hydrogen}}{\text{total formula mass}} \times 100\% = \frac{18.18\ u}{114.26\ u} \times 100\% = 15.91\%$$

How chemists measure formula mass and percent composition

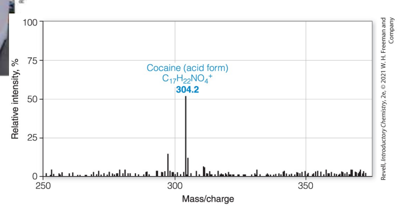


Mass Spectrometry

a technique used to measure the mass of molecules



Mass Spectrometry, Continued



Elemental Analysis

a technique used to measure percent composition

uses combustion reactions to form simpler products (CO_2, H_2O)

a b

The Mole Concept

1 atomic mass unit (u) = 1.66×10^{-24} g

How do we relate atomic masses to larger amounts?

The Mole Concept, Continued

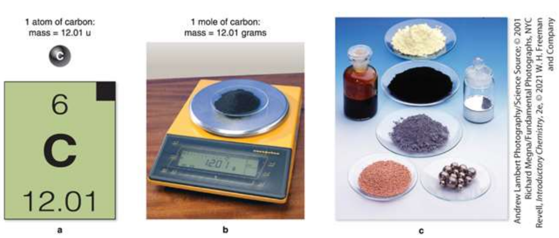
Avogadro's Number: 6.02×10^{23}

1 dozen: 12 units
1 dozen planets = 12 planets
1 dozen toothpicks = 12 toothpicks
1 dozen donuts = 12 donuts

1 mole: 6.02×10^{23} units
1 mole of donuts = 6.02×10^{23} donuts
1 mole of carbon atoms = 6.02×10^{23} carbon atoms
1 mole of oxygen molecules = 6.02×10^{23} oxygen molecules

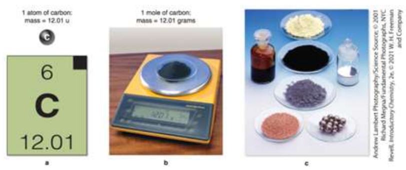
Moles relate atoms to grams, Part 1

- 1 atom of carbon = 12.01 u
- 1 mole of carbon = 12.01 g



Moles relate atoms to grams, Part 2

- 1 atom of carbon = 12.01 u
- 1 mole of carbon = 12.01 g



Mass of carbon:
12.01 u
12.01 grams/mole
(molar mass)

Moles relate atoms to grams, Part 3

6 C 12.01
8 O 16.00

What is the formula mass of carbon dioxide?

- 1 molecule of CO_2 = 44.01 u
- 1 mole of CO_2 = 44.01 g

Converting between Grams and Moles

Use molar mass as the conversion factor

How many moles of NaCl are present in a 305-gram sample?

formula mass of NaCl: 58.44 g/mole

58.44 g NaCl = 1 mole NaCl

$$305 \text{ g NaCl} \times \frac{1 \text{ mole NaCl}}{58.44 \text{ g NaCl}} = 5.22 \text{ moles NaCl}$$

Converting between Grams and Moles, Continued

Use molar mass as the conversion factor

To prepare a solution that contains 1.20 moles of NaCl, how many grams of NaCl are needed?

$$1.20 \text{ ~~moles NaCl~~} \times \frac{58.44 \text{ g NaCl}}{1 \text{ ~~mole NaCl~~}} = 70.1 \text{ g NaCl}$$

Converting between Moles and Particles

6.02×10^{23} particles = 1 mole

How many atoms are in 4.20 moles of gold?

$$4.20 \text{ ~~moles~~ Au} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ ~~mole~~}} = 2.53 \times 10^{24} \text{ atoms Au}$$

Relating Atoms to Grams

What is the mass in grams of 2.53×10^{23} iron atoms?

55.85 g Fe = 1 mole Fe

6.02×10^{23} atoms = 1 mole

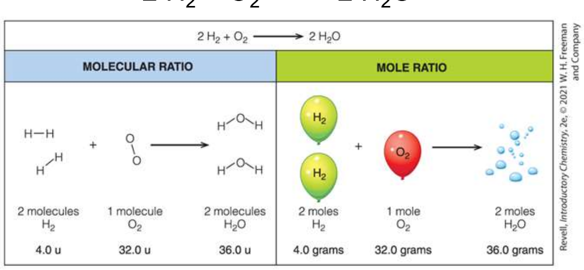
$$2.53 \times 10^{23} \text{ ~~atoms Fe~~} \times \frac{1 \text{ ~~mole Fe~~}}{6.02 \times 10^{23} \text{ ~~atoms Fe~~}} \times \frac{55.85 \text{ g Fe}}{1 \text{ ~~mole Fe~~}} = 23.5 \text{ grams Fe}$$

Relating Grams to Atoms or Molecules

The Mole Concept in Balanced Equations

Equation Coefficients Can Mean Molecules or Moles

$$2\ H_2 + O_2 \longrightarrow 2\ H_2O$$



Using the Mole Concept

If 235 grams of iron react in this way…

$Fe_{(s)} + 2\,HCl_{(aq)} \rightarrow FeCl_{2\,(aq)} + H_{2\,(g)}$

4.21 moles   8.42 moles   4.21 moles

How many moles of iron react?	How many moles of HCl are needed?

Fe: 55.85 g/mol

$235\ \text{g Fe} \times \frac{1\ \text{mol Fe}}{55.85\ \text{g Fe}} = 4.21\ \text{mol Fe}$

$4.21\ \text{mol Fe} \times \frac{2\ \text{mol HCl}}{1\ \text{mol Fe}} = 8.42\ \text{mol HCl}$

How many moles of iron(II) chloride form?	How many grams of iron(II) chloride form?

$4.21\ \text{mol Fe} \times \frac{1\ \text{mol } FeCl_2}{1\ \text{mol Fe}} = 4.21\ \text{mol } FeCl_2$

126.75 g/mol

$4.21\ \text{mol } FeCl_2 \times \frac{126.75\ \text{g } FeCl_2}{1\ \text{mol } FeCl_2} = 534\ \text{g } FeCl_2$

Using the Mole Concept, Continued

If 235 grams of iron react in this way…

$Fe_{(s)} + 2\,HCl_{(aq)} \rightarrow FeCl_{2\,(aq)} + H_{2\,(g)}$

How many moles of iron react?	How many moles of HCl are needed?
How many moles of iron(II) chloride form?	How many grams of iron(II) chloride form?

Stoichiometry Using the amount of one material to predict the amount of another, based on the balanced equation.

Using the Mole Concept, Practice

When magnesium burns, it combines with oxygen to form MgO. If this reaction consumes 3.0 moles of oxygen, how many moles of MgO will form? How many grams of MgO will form?

$2\,Mg_{(s)} + O_{2\,(g)} \rightarrow 2\,MgO_{(s)}$

3.0 mol   6.0 mol

$3.0\ \text{mol } O_2 \times \frac{2\ \text{mol MgO}}{1\ \text{mol } O_2} = 6.0\ \text{mol MgO}$

MgO: 40.30 g/mol

$6.0\ \text{mol MgO} \times \frac{40.30\ \text{g MgO}}{1\ \text{mol MgO}} = 240\ \text{g MgO}$

Using the Mole Concept, More Practice

Sodium metal reacts violently with water. How many moles of H_2 gas form if 11.0 grams of sodium react with water?

$$2\ Na_{(s)} + 2\ H_2O_{(l)} \rightarrow 2\ NaOH_{(aq)} + H_{2\ (g)}$$

22.99 g/mol

g Na ⇨ mol Na $\quad 11.0\ \cancel{g\ Na} \times \frac{1\ mol\ Na}{22.99\ \cancel{g\ Na}} = 0.478\ mol\ Na$

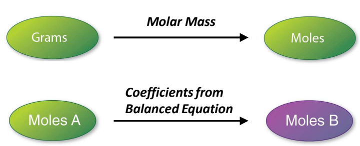
mol Na ⇨ mol H_2 $\quad 0.478\ \cancel{mol\ Na} \times \frac{1\ mol\ H_2}{2\ \cancel{mol\ Na}} = 0.239\ mol\ H_2$

$$11.0\ \cancel{g\ Na} \times \frac{1\ \cancel{mol\ Na}}{22.99\ \cancel{g\ Na}} \times \frac{1\ mol\ H_2}{2\ \cancel{mol\ Na}} = 0.239\ mol\ H_2$$

Summary of Stoichiometry Problems

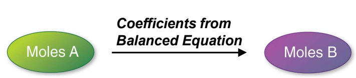
$$Fe_{(s)} + 2\ HCl_{(aq)} \rightarrow FeCl_{2\ (aq)} + H_{2\ (g)}$$

Conversion Factors in Stoichiometry Problems:

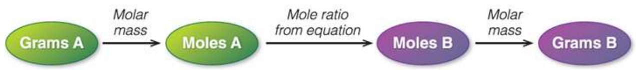


The Mole Concept in Balanced Equations

$$Fe_{(s)} + 2\ HCl_{(aq)} \rightarrow FeCl_{2\ (aq)} + H_{2\ (g)}$$



Gram-to-Gram Problems



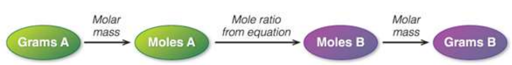
Gram-to-Gram Problems, Practice

When heated with a Bunsen burner, $MgCO_3$ decomposes to MgO and CO_2, as shown in this equation. If 5.24 g of $MgCO_3$ are heated in this manner, how many grams of MgO can be produced?

$$MgCO_3 \rightarrow MgO + CO_2$$

$$g\ MgCO_3 \rightarrow mol\ MgCO_3 \rightarrow mol\ MgO \rightarrow g\ MgO$$

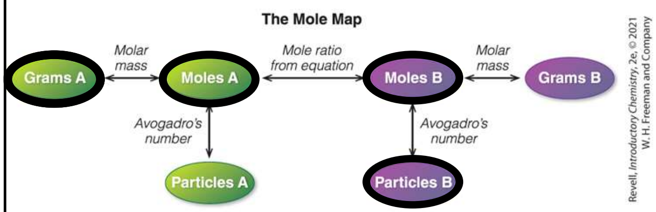
$$5.24\ \cancel{g\ MgCO_3} \times \frac{1\ \cancel{mol\ MgCO_3}}{84.32\ \cancel{g\ MgCO_3}} \times \frac{1\ \cancel{mol\ MgO}}{1\ \cancel{mol\ MgCO_3}} \times \frac{40.31\ g\ MgO}{1\ \cancel{mol\ MgO}} = 2.51\ g\ MgO$$



Strategies for Solving Stoichiometry Problems

Conversion Type	Conversion Factor
Grams and moles of one substance	Molar Mass
Moles and particles of one substance	Avogadro’s number
Moles of two different substances	Mole ratio from the balanced equation

The Mole Map

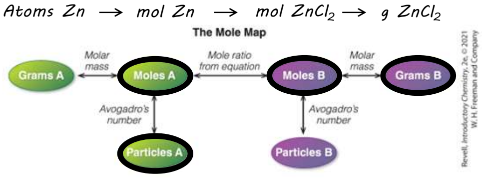


Using the Mole Map, Practice

Zinc metal reacts with aqueous copper(II) chloride, as shown in this equation. If 3.03×10^{21} atoms of zinc react, how many grams of $ZnCl_2$ will form? Show the sequence of conversions necessary, then calculate the numerical answer.

$$Zn\ (s) + CuCl_2\ (aq) \rightarrow ZnCl_2\ (aq) + Cu\ (s)$$

Atoms Zn → mol Zn → mol $ZnCl_2$ → g $ZnCl_2$



Using the Mole Map, Practice Continued

Zinc metal reacts with aqueous copper(II) chloride, as shown in this equation. If 3.03×10^{21} atoms of zinc react, how many grams of $ZnCl_2$ will form? Show the sequence of conversions necessary; then calculate the numerical answer.

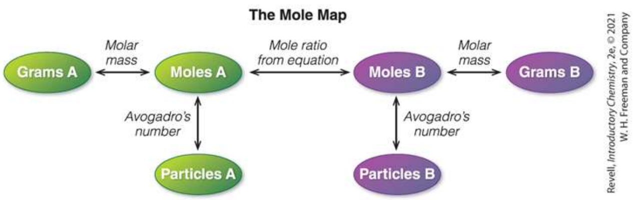
$$Zn\ (s) + CuCl_2\ (aq) \rightarrow ZnCl_2\ (aq) + Cu\ (s)$$

Atoms Zn → mol Zn → mol $ZnCl_2$ → g $ZnCl_2$

$$3.03 \times 10^{21}\ \cancel{\text{atoms Zn}} \times \frac{1\ \cancel{\text{mol Zn}}}{6.02 \times 10^{23}\ \cancel{\text{atoms Zn}}} \times \frac{1\ \cancel{\text{mol } ZnCl_2}}{1\ \cancel{\text{mol Zn}}} \times \frac{136.28\ \text{g } ZnCl_2}{1\ \cancel{\text{mol } ZnCl_2}}$$

$$= 0.686\ \text{g } ZnCl_2$$

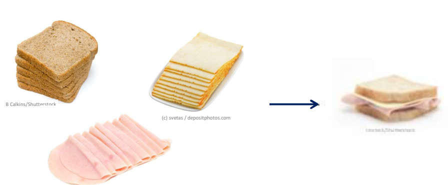
Using the Mole Map, Summary



The Mole Concept in Balanced Equations: Limiting Reagents

$$Fe_{(s)} + 2\,HCl_{(aq)} \rightarrow FeCl_{2\,(aq)} + H_{2\,(g)}$$

Calculations with Limiting Reagents



2 slices bread + 1 slice turkey + 1 slice cheese → 1 sandwich

Calculations with Limiting Reagents, Practice

If you have 80 slices of bread, 18 slices of turkey, and 15 slices of cheese, how many turkey-and-cheese sandwiches can you make using this recipe?

2 slices bread + 1 slice turkey + 1 slice cheese → 1 sandwich

Bread: 40 sandwiches

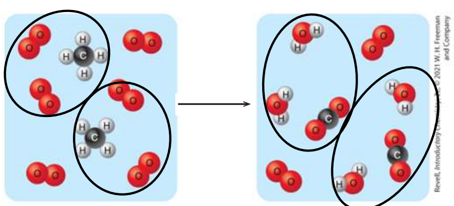
Turkey: 18 sandwiches

Cheese: 15 sandwiches

Limiting Reagent: runs out first
limits the amount that can be produced

Limiting Reagent Reactions Are Common

$CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (g)$



CH_4 is the limiting reagent.
O_2 is the excess reagent.

Limiting Reagents, Practice

Potassium reacts violently with chlorine gas to produce potassium chloride, as shown. If 1.2 moles of potassium are combined with 15 moles of chlorine gas, how many moles of potassium chloride can form? Which reagent is the limiting reagent?

$2\ K\ (s) + Cl_2\ (g) \rightarrow 2\ KCl\ (s)$

1.2 moles (K) 15 moles (Cl_2)

We have enough K to produce 1.2 moles of KCl:

$1.2\ \text{mol K} \times \frac{2\ \text{mol KCl}}{2\ \text{mol K}} = 1.2\ \text{moles KCl}$

K is the limiting reagent.
1.2 moles of KCl can be produced.

We have enough Cl_2 to produce 30 moles of KCl:

$15\ \text{mol}\ Cl_2 \times \frac{2\ \text{mol KCl}}{1\ \text{mol}\ Cl_2} = 30\ \text{moles KCl}$

Cl_2 is the excess reagent.
There will be Cl_2 left over after the reaction is complete.

Limiting Reagents, More Practice

Uranium reacts with fluorine gas according to this equation. If 30 moles of uranium combine with 75 moles of F_2, how many moles of UF_6 will form?

$$U + 3\,F_2 \rightarrow UF_6$$

30 ~~mol U~~ x $\frac{1 \text{ mol } UF_6}{1 \text{ ~~mol U~~}}$ = 30 moles UF_6 — U is the excess reagent.

75 ~~mol F_2~~ x $\frac{1 \text{ mol } UF_6}{3 \text{ ~~mol } F_2\text{~~}}$ = 25 moles UF_6 — F_2 is the limiting reagent.

The ICE Method

If you have 80 slices of bread, 18 slices of turkey, and 15 slices of cheese, how many turkey-and-cheese sandwiches can you make using this recipe?

If you make all of the sandwiches, what will be left over?

	2 slices bread	+ 1 slice turkey	+ 1 slice cheese	→ 1 sandwich
I nitial	80	18	15	0
C hange	-30	-15	-15	+15
E nd	50	3	0	15

The ICE Method Practice

If 15 moles of HCl and 20 moles of NaOH are combined, how many moles of each species will be present after the reaction is complete?

$$HCl + NaOH \rightarrow NaCl + H_2O$$

	HCl	NaOH	NaCl	H_2O
I nitial	15 mol	20 mol	0 mol	0 mol
C hange	-15 mol	-15 mol	+15 mol	+15 mol
E nd	0 mol	5 mol	15 mol	15 mol

Summary of Limiting Reagents

- Limiting Reagent: Completely consumed; limits the amount of product formed.
  - The reagent that forms the least amount of product is the limiting reagent.
- Excess Reagent: Not completely consumed; reagent will be left over after the reaction is complete.
- ICE method: Can be used to determine the amounts of all reactants and products present after a reaction.

Theoretical and Percent Yield

- **Theoretical Yield:** The amount of a product that can form, based on the balanced equation.
- **Actual Yield:** The amount actually obtained.
- **Percent Yield:** The percentage of the theoretical yield that was obtained.

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

Why is the Actual Yield so Low?

- Material sticks to container walls
- Unwanted side products
- Product lost during purification

Percent Yield Practice

A chemist runs a reaction in which the theoretical yield is 240 grams. However, he is only able to isolate 180 grams. What is the percent yield for this reaction?

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

$$= \frac{180 \text{ g}}{240 \text{ g}} \times 100\% = 75\%$$

Percent Yield, More Practice

Sulfur hexafluoride, SF_6, is widely used in the power industry. It is produced through this reaction:

$$S_{(s)} + 3\,F_{2\,(g)} \rightarrow SF_{6\,(g)}$$

A manufacturer reacts 120.0 kilograms of sulfur with excess fluorine gas. What mass of SF_6 is theoretically possible for this conversion? After the reaction is complete, the manufacturer isolates 480.2 kilograms of SF_6. What was the percent yield for this process?

$$\text{g S} \rightarrow \text{mol S} \rightarrow \text{mol } SF_6 \rightarrow \text{g } SF_6$$

$$120{,}000 \text{ g S} \times \frac{1 \text{ mol S}}{32.06 \text{ g S}} \times \frac{1 \text{ mol } SF_6}{1 \text{ mol S}} \times \frac{146.06 \text{ g } SF_6}{1 \text{ mol } SF_6} = 546{,}700 \text{ g } SF_6$$

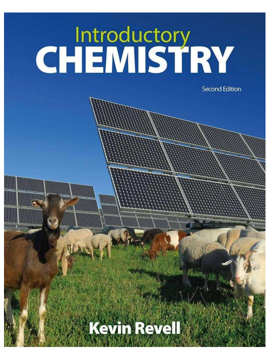
$$= 546.7 \text{ kg } SF_6$$

$$\% \text{ yield} = \frac{480.2 \text{ kg}}{546.7 \text{ kg}} \times 100\% = 87.84\%$$

Introductory Chemistry
Chem 103

Chapter 8 – Energy

Lecture Slides



Energy, Work, and Heat

Thermodynamics: the study of energy and temperature changes

Thermochemistry: energy changes in chemical reactions

Energy: the ability to do work

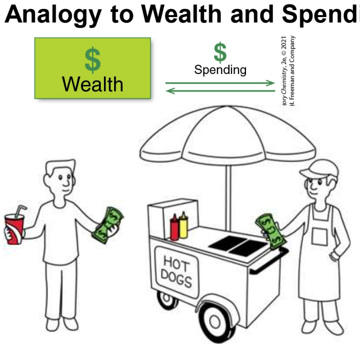
Forms of energy:

- **potential energy:** stored energy
- **kinetic energy:** energy of motion

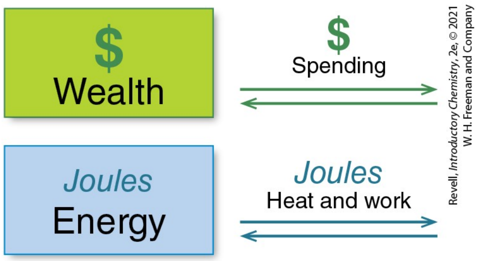
Types of energy changes:

- **heat:** the transfer of kinetic energy
- **work:** the transfer of energy from one form to another.

An Analogy to Wealth and Spending

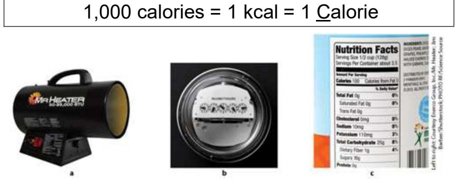


An Analogy to Wealth and Spending, Continued



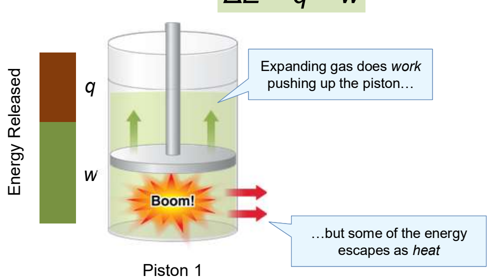
Units of Energy

1 joule (J) = 1 kg·m^2/s^2
1 British Thermal Unit (BTU) = 1,055 J
1 kilowatt-hour (kWh) = 3.6 × 10^6 J
1 calorie (c) = 4.184 J
1,000 calories = 1 kcal = 1 Calorie



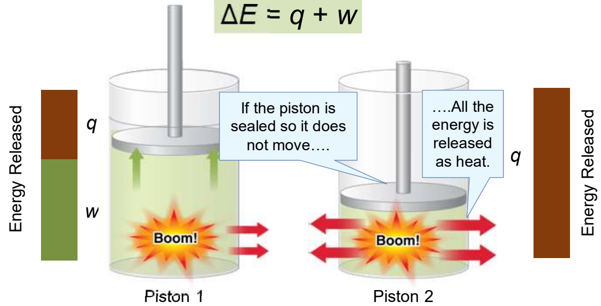
The Relationship Between Energy, Heat, and Work

$$\Delta E = q + w$$



Energy, Heat, and Work, Continued

$\Delta E = q + w$



Energy, Heat, and Work Practice

A small sample of propane burns, producing carbon dioxide and water vapor. As the hot gas mixture expands, it releases 20.0 kJ of heat, and does 31.0 kJ of work pushing against a piston. What is the total amount of energy released in this reaction?

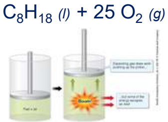
Energy released = q + w

= 20.0 kJ + 31.0 kJ

= 51.0 kJ

Endothermic and Exothermic Changes

$2\ C_8H_{18\ (l)} + 25\ O_{2\ (g)} \rightarrow 16\ CO_{2\ (g)} + 18\ H_2O_{\ (g)}$ + energy



exothermic change
releases energy

energy + $H_2O_{\ (s)} \rightarrow H_2O_{\ (l)}$

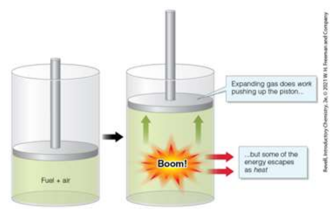
Photo credit: Zoom Team/Shutterstock

endothermic change
absorbs energy

Identifying the System and Surroundings

System: the part of the universe being studied

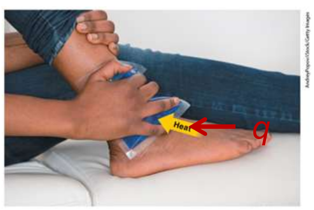
Surroundings: the rest of the universe



Identifying the System and Surroundings, Continued

System: the part of the universe being studied

Surroundings: the rest of the universe



Showing the direction of energy changes

Exothermic change

releases heat energy

→ $-q$

System does work on surroundings

releases energy

$-w$

Endothermic change

absorbs heat energy

← $+q$

Surroundings does work on system

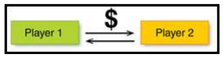
absorbs energy

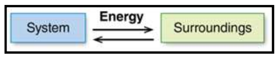
$+w$

The Law of Conservation of Energy

Energy cannot be created or destroyed.

$$\Delta E_{system} = -\Delta E_{surroundings}$$





The Law of Conservation of Energy Practice

A chemical reaction releases 200 J of heat energy to its surroundings. Write this change of energy for the system (the chemical reaction), and for the surroundings.

System: $\Delta E = -200$ J

Surroundings: $\Delta E = +200$ J

Summary of Energy Changes

- Energy changes: work and heat
  - System
  - Surroundings
  - Exothermic reaction: system releases heat
  - Endothermic reaction: system absorbs heat
- Energy is not created or destroyed in chemical reactions.

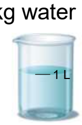
Heat Energy and Temperature

Heat The total kinetic energy transferred from one substance or object to another.

Temperature The average kinetic energy of the particles in a substance.

Specific Heat and Heat Capacity

1 kg water —1,000 calories of heat energy→ Temp ↑ 1.0 °C



1 kg iron —1,000 calories of heat energy→ Temp ↑ 9.3 °C

Specific Heat

Specific heat: The amount of heat required to raise the temperature of 1 gram of material by 1 °C.

$$\text{specific heat} = \frac{\text{heat}}{(\text{mass}) \times (\text{change in temperature})}$$

$$s = \frac{q}{m\Delta T}$$

$$q = ms\Delta T$$

Different Materials have Different Specific Heats

TABLE 8.2 Specific Heats for Several Materials

	Material	Specific Heat (J/g ·°C)
Gas	Air (dry)	1.01
Liquid	Water (liquid)	4.184
	Ethanol	2.597
	Oil (petroleum)	1.74
	Gasoline	2.2
Solid	Glass (quartz)	0.70
	Concrete	0.880
	Ice	2.10
	Sand	0.799
	Aluminum	0.897
	Chromium	0.449
	Gold	0.129
	Iron	0.449
	Lead	0.130
	Nickel	0.444
	Zinc	0.388
	Steel	0.50


California CPA/Moment Open/ Getty Images

Specific Heat Calculations Practice

How many kilojoules of heat are required to raise the temperature of 120.0 grams of water by 5.0 °C?

$q = ms\Delta T$

$= (120.0\ \text{g})(4.184\ \text{J/(g·°C)})(5.0\ \text{°C})$

$= 2{,}500\ \text{J}$

$= 2.5\ \text{kJ}$

Comparing Specific Heat and Heat Capacity

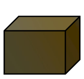
specific heat, *s*

The amount of heat required to raise **1 g** by 1 °C.

$$\text{specific heat} = \frac{q}{m\Delta T}$$

$$s = \frac{q}{m\Delta T}$$

$$q = ms\Delta T$$



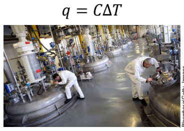
heat capacity, *C*

The amount of heat required to raise an object by 1 °C.

$$\text{heat capacity} = \frac{q}{\Delta T}$$

$$C = \frac{q}{\Delta T}$$

$$q = C\Delta T$$


Bloomberg/Getty Images

Heat Capacity Calculations Practice

When filled with water, a large reaction vessel in a chemical plant has a heat capacity of 5.41 × 10^5 kJ/°C. How many kJ of heat are required to heat this entire vessel from 25.0 °C to 48.2 °C?

$\Delta T = T_{final} - T_{initial}$

$= 48.2\ °C - 25.0\ °C = 23.2\ °C$

$q = C\Delta T$

$= (5.41 \times 10^5\ kJ/\cancel{°C})(23.2\ \cancel{°C})$

$= 1.26 \times 10^7\ kJ$

Calorimetry

Calorimetry experiments – measure the flow of heat

coffee cup calorimetry

bomb calorimetry

Coffee-Cup Calorimetry

Fe

$q_{water} = -q_{metal}$

$q = ms\Delta T$

(m: mass; s: specific heat; ΔT: change in temp.)

$m_w s_w \Delta T_w = -m_m s_m \Delta T_m$

$$s_m = -\frac{m_w s_w \Delta T_w}{m_m \Delta T_w}$$

Coffee-cup Calorimetry Calculations Practice

A chemist heats a 26.0-g sample of an unknown metal to 100.0 °C, then places it in a coffee-cup calorimeter containing 52.1 g of water at an initial temperature of 20.0 °C. After some time, both the metal and water reach an equal temperature of 24.0 °C. What is the specific heat of the metal? (s_w = 4.184 J/g·°C)

s_w = 4.184 J/g·°C

m_w = 52.1 g

m_m = 26.0 g

$\Delta T_m = T_{final} - T_{initial}$

= 24.0 °C − 100.0 °C

= −76.0 °C

$\Delta T_w = T_{final} - T_{initial}$

= 24.0 °C − 20.0 °C

= 4.0 °C

$q_{water} = -q_{metal}$

$$s_m = \frac{-m_w s_w \Delta T_w}{m_m \Delta T_m}$$

$$s_m = \frac{-(52.1\ g)(4.184\ J/(g\cdot °C))(4.0\ °C)}{(26.0\ g)(-76.0\ °C)}$$

s_m = 0.44 J/g·°C

Coffee-cup Calorimetry Calculations, Continued

heat + $NH_4Cl_{(s)}$ → $NH_4Cl_{(aq)}$

$q_{solid} = -q_{aq}$ $q = ms\Delta T$

$q_{solid} = -m_{aq}s_{aq}\Delta T_{aq}$

Photo credits: Martyn F. Chillmaid/Science Source

Coffee-cup Calorimetry, More Practice

A 10.4-gram sample of NH_4Cl was combined with 100.0 grams of water in a coffee-cup calorimeter, causing the water temperature to decrease by 6.20 °C. Based on this, how much heat energy was required to dissolve the sample of NH_4Cl? Calculate the heat of solution for NH_4Cl in kJ/mol.

m_{aq} = 10.4 g + 100.0 g

= 110.4 g

s_{aq} = 4.184 J/g·°C

ΔT_{aq} = − 6.20 °C

$q_{solid} = -m_{aq}s_{aq}\Delta T_{aq}$

q_{solid} = −(110.4 g)(4.184 J/(g·°C))(−6.20 °C)

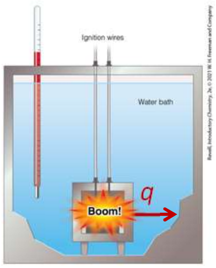
= 2,860 J = 2.86 kJ

$$10.4\ g\ NH_4Cl \times \frac{1\ mole\ NH_4Cl}{53.49\ g\ NH_4Cl} = 0.194\ moles\ NH_4Cl$$

$$\text{Heat of solution } (NH_4Cl) = \frac{2.86\ kJ}{0.194\ mol} = 14.7\ kJ/mol$$

Bomb Calorimetry

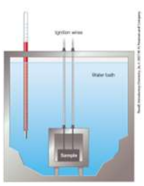
measures energy content (heats of reaction) in food and fuels



$q = C\Delta T$

Bomb Calorimetry Calculations Practice

A chemist places a 20.0-g sample of ethanol inside a bomb calorimeter with a known heat capacity of 28.72 kJ/°C. When the ethanol ignites, the temperature of the calorimeter rises from 22.04 °C to 42.74 °C. How much heat did the ethanol release? Calculate the energy released in kilojoules per gram of ethanol.



$\Delta T = 20.70\ °C$

$C = 28.72\ kJ/°C$

$q = C\Delta T$

$q = (28.72\ kJ/°C)(20.70\ °C)$

$= 594.5\ kJ$

$\frac{594.5\ kJ}{20.0\ g} = 29.7\ kJ/g$

Heat Energy and Chemical Reactions

Chemical reactions involve changes in energy.

The energy of a reaction is an extensive property.

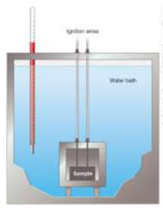
Fuel Value

Fuel value The amount of energy that can be produced by the combustion of a material

TABLE 8.3 Fuel Values for Common Combustion Fuels

Fuel	Fuel Value (kJ/g)
Methane	55.5
Natural gas	54.0
Propane	50.3
Butane	49.5
Gasoline	46.5
Anthracite coal	34.6
Ethanol	29.7
Wood (oak)	18.9

Data from *CRC Handbook of Chemistry and Physics*, 92nd ed. (Boca Raton, FL: CRC Press, 2011).



Reaction Enthalpy, ΔH_{rxn}

Reaction enthalpy The amount of heat energy absorbed or released in a chemical reaction at constant pressure.

$C_2H_6O + 3\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O$ $\Delta H_{rxn} = -1{,}368$ kJ

exothermic

$\frac{-1{,}368 \text{ kJ}}{1 \text{ mol } C_2H_6O}$ *or* $\frac{-1{,}368 \text{ kJ}}{3 \text{ mol } O_2}$ *or* $\frac{-1{,}368 \text{ kJ}}{2 \text{ mol } CO_2}$ *or* $\frac{-1{,}368 \text{ kJ}}{3 \text{ mol } H_2O}$

Reaction Enthalpy Calculations Practice

How much heat will be released by the combustion of 789.0 g of ethanol, C_2H_6O?

$C_2H_6O + 3\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O$ $\Delta H_{rxn} = -1{,}368$ kJ

Grams → Moles → Energy

$$789.0 \text{ g } C_2H_6O \times \frac{1 \text{ mole } C_2H_6O}{46.08 \text{ g } C_2H_6O} \times \frac{-1{,}368 \text{ kJ}}{1 \text{ mole } C_2H_6O} = -2.342 \times 10^4 \text{ kJ}$$

2.342×10^4 kJ of heat energy released

Reaction Enthalpy Calculations, More Practice

Many manufacturers produce hydrogen gas from methane gas, as shown in the reaction below. This reaction is endothermic, with a ΔH_{rxn} = 206.1 kJ. How much heat energy is required to produce 1.00 kg of hydrogen gas?

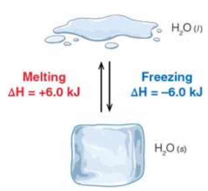
$CH_4(g) + H_2O(g) \rightarrow CO(g) + 3\ H_2(g)$ $\Delta H_{rxn} = 206.1$ kJ

$$1.00\ \text{kg}\ H_2 \times \frac{1{,}000\ \text{g}}{1\ \text{kg}} \times \frac{1\ \text{mol}\ H_2}{2.02\ \text{g}\ H_2} \times \frac{206.1\ \text{kJ}}{3\ \text{mol}\ H_2} = 3.40 \times 10^4\ \text{kJ}$$

Physical Changes Involve Enthalpy Changes

Melting: $H_2O\ (s) \rightarrow H_2O\ (l)$ $\Delta H = 44.0$ kJ

Freezing: $H_2O\ (l) \rightarrow H_2O\ (s)$ $\Delta H = -44.0$ kJ



Summary of Energy

Reaction energy is an extensive property.

Fuel value is the energy released in a combustion reaction.

The reaction enthalpy relates heat released in a reaction to the balanced equation.