

# UNIT (1) MEASUREMENTS IN CHEMISTRY

Measurements are part of our daily lives. We measure our weights, driving distances, and gallons of gasoline. As a health professional you might measure blood pressure, temperature, pulse rate, drug dosage, or percentage of body fat.

A **measurement** contains a *number* and a *unit*.

A *unit* specifies the physical property and the size of a measurement, while the *number* indicates how many units are present. A number without a unit is usually meaningless.

## 1.1 | Units of Measurement

In the United States most measurements are made with the English system of units which usually contain fractions (a collection of functionally unrelated units.)

The **metric system** is a decimal-based system of units of measurement which is used most often worldwide.

Around 1960, the international scientific organization adopted a modification of the metric system called **International System** or **SI** (from *Système International*).

Quantity	English Unit	Metric Unit	SI Unit
Mass	pound (lb)	gram (g)	kilogram (kg)
Length	foot (ft)	meter (m)	meter (m)
Volume	quart (qt)	liter (L)	cubic meter (m <sup>3</sup> )
Temperature	degree Fahrenheit (°F)	degree Celsius (°C)	Kelvin (K)
Energy	calorie (cal)	calorie (cal)	Joule (J)

## 1.2 | Scientific Notation

**Scientific notation** is a common method used to conveniently represent very small or very large numbers. There are two parts to any number expressed in scientific notation, a coefficient, and a power of 10. The number 683 is written in scientific notation as  $6.83 \times 10^2$ . The coefficient is 6.83 and  $10^2$  shows the power of 10 (the superscript 2 is called an exponent). A number less than one would contain a negative exponent. For example: the number 0.0075 is written as  $7.5 \times 10^{-3}$  (note the negative exponent).

The coefficient must always be a number greater than or equal to 1 but less than 10 or  $1 \leq \text{coefficient} < 10$ .

**Worked Example 1-1**

Express the following numbers in scientific notation:

- a) 408.00      b) 0.007956

**Solution**

Apply the following:

Place the decimal point after the first nonzero digit in the number.

Indicate the number of places the decimal was moved using the power of 10.

If the decimal is moved to the left, the power of 10 is positive. If moved to the right, it is negative.

a)  $4.0800 \times 10^2$  (coefficient = 4.0800, exponent = +2)

b)  $7.956 \times 10^{-3}$  (coefficient = 7.956, exponent = -3)

**Practice 1-1**

Express each of the following values in scientific notation:

a) There are 33,000,000,000,000,000 molecules of water in one milligram of water.

b) A single molecule of sucrose weighs 0.000 000 000 000 000 000 57 g.

**Answer****Practice 1-2**

Convert each the following scientific notation to decimal notation.

a)  $8.54 \times 10^3$

b)  $6.7 \times 10^{-5}$

c)  $1.29 \times 10^4$

d)  $1.000 \times 10^{-2}$

**Answer**

## Scientific Notation and Calculators

Numbers in scientific notation can be entered into most calculators using the EE or EXP key. As an example try  $9.7 \times 10^3$ .

1. Enter the coefficient (9.7) into calculator.
2. Push the EE (or EXP) key. Do NOT use the x (times) button.
3. Enter the exponent number (3).

Number to Enter	Method	Display Reads
$9.7 \times 10^3$	9.7 EE or EXP 3	$9.7^{03}$ or 9.7E03 or 9700

Now try  $8.1 \times 10^{-5}$ :

1. Enter the coefficient (8.1) into calculator.
2. Push the EE (or EXP) key. Do NOT use the x (times) button.
3. Enter the exponent number (5). Use the plus/minus (+/-) key to change its sign.

Number to Enter	Method	Display Reads
$8.1 \times 10^{-5}$	8.1 EE or EXP 5 +/-	$8.1^{-05}$ or 8.1E-05

## 1.3 | Metric Prefixes

The **metric system** is a decimal-based system of units of measurement used by most scientists worldwide.

In the metric system, a prefix can be attached to a unit to *increase* or *decrease* its size by factors (powers) of 10.

	Prefix	Value
↑	tera- (T)	$10^{12} = 1,000,000,000,000$
	giga- (G)	$10^9 = 1,000,000,000$
	mega- (M)	$10^6 = 1,000,000$
	kilo- (k)	$10^3 = 1,000$
↓	deci- (d)	$10^{-1} = 0.1$
	centi- (c)	$10^{-2} = 0.01$
	milli- (m)	$10^{-3} = 0.001$
	micro- ( $\mu$ )	$10^{-6} = 0.000001$
	nano- (n)	$10^{-9} = 0.000000001$
	pico- (p)	$10^{-12} = 0.000000000001$

### Practice 1-3

Give the metric prefix that corresponds to each of the following

- a) 1,000,000,000   b)  $10^{-6}$    c) 1000   d) 0.01   e)  $10^{-9}$    f)  $10^{12}$

Answer

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## 1.4 | Significant Figures in Measurements

A student is asked to determine the mass of a small object using two different balances available in the lab. The lower priced model reports masses to within  $\pm 0.01$  g (one-one hundredth), while the higher priced one reports to within  $\pm 0.0001$  g (one-ten thousandth). The student measures the mass three times on each balance and completes the following table.

	First balance	Second balance
Three measurements	<b>2.16, 2.14, 2.15 g</b>	<b>2.1538, 2.1539, 2.1537 g</b>
Average	<b>2.15 g</b>	<b>2.1538 g</b>
Reproducibility	<b><math>\pm 0.01</math> g</b>	<b><math>\pm 0.0001</math> g</b>
Which digit is the “uncertain digit” in the average?	<b>The last digit; 5</b>	<b>The last digit; 8</b>
Which digits are “certain digits” in the average?	<b>2, 1</b>	<b>2, 1, 5, 3</b>
How many significant digits are in the average?	<b>Three significant digits</b>	<b>Five significant digits</b>

**Significant figures (sig figs)** are the digits that are known with certainty plus one digit that is uncertain. **All** nonzero digits in measurements are always significant.

### Are zeroes significant?

**YES:** zeros between nonzero digits (20509).

**YES:** zeros at the end of a number when a decimal point is written (3600.).

**NO:** zeros at the end of a number when no decimal point is written (3600).

**NO:** zeros at the beginning of a number (0.0047).

### Worked Example 1-2

How many significant figures does each number have?

- a) 0.0037                      b) 600.                      c) 93,000  
d)  $2.08 \times 10^{-5}$               e) 600                      f) 58.00  
g) 4010049                      h)  $1.700 \times 10^2$               i)  $4.0100 \times 10^6$

#### Solution

	sf		sf		sf
0.0037	<b>2</b>	600.	<b>3</b>	93,000	<b>2</b>
$2.08 \times 10^{-5}$	<b>3</b>	600	<b>1</b>	58.00	<b>4</b>
4010049	<b>7</b>	$1.700 \times 10^2$	<b>4</b>	$4.0100 \times 10^6$	<b>5</b>

### Significant Figures in “Exact Numbers”

Exact numbers have an **unlimited** number of significant figures. Exact numbers are obtained by **counting** items or by **definition**.

Counting: 24 students mean 24.0000000... students. 8 pennies means 8.0000... pennies.

Definition: 1 m = 100 cm means 1.00000...m = 100.000000...cm

## 1.5 | Calculations Involving Significant Figures

### Rules for Rounding off Numbers

If the first digit to be deleted is 4 or less, leave the last reported digit unchanged.

If the first digit to be deleted is 5 or greater, increase the last reported digit by one.

In some cases you need to *add* significant zeros. The number 2, reported in four significant figures, is 2.000.

### Practice 1-4

Round off each of the following to three significant figures.

- a) 9.174                      b) 9.175                      c) 9.176  
d) 5                          e) 0.0040                      f) 8000  
g)  $2.4 \times 10^{-5}$                       h) 670

#### Answer

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## Rules for Rounding off in Calculations

### A. Multiplication and Division

The answer carries the **same number of significant figures** as the factor with the fewest significant figures.

#### Practice 1-5

Perform each of the following calculations to the correct number of significant figures.

a)  $33.56 \times 1.9483$

b)  $(2.50 \times 10^{-3}) \times (1.8500 \times 10^5)$

c)  $47.5301 \div 2.30$

d)  $(6.56 \times 10^{10}) \div (7.8 \times 10^9)$

**Answer**

### B. Addition and Subtraction

The answer should have the **same number of decimal places** as the quantity with the fewest decimal places.

#### Practice 1-6

Perform each of the following calculations to the correct number of significant figures:

a)  $73.498 + 2.2$

b)  $63.81 + 205.4$

c)  $191.000 - 188.0$

d)  $124.08 - 39.1740$

e)  $(6.8 \times 10^{-2}) + (2.04 \times 10^{-2})$

f)  $(5.77 \times 10^{-4}) - (3.6 \times 10^{-4})$

**Answer**

## 1.6 | Writing Conversion Factors

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Many problems in chemistry require converting a quantity from one unit to another. To perform this conversion, you must use a **conversion factor** or series of conversion factors that relate two units. This method is called **dimensional analysis**.

Any equality can be written in the form of a fraction called a conversion factor. A conversion factor is easily distinguished from all other numbers because it is always a fraction that contains different units in the numerator and denominator.

Converting kilograms to pounds can be performed using the equality  $1 \text{ kg} = 2.20 \text{ lb}$ . The two different conversion factors that may be written for the equality are shown below. Note the different units in the numerator and denominator, a requirement for all conversion factors.

**Conversion Factors:**  $\frac{\text{Numerator}}{\text{Denominator}}$      $\frac{1 \text{ kg}}{2.20 \text{ lb}}$     or     $\frac{2.20 \text{ lb}}{1 \text{ kg}}$

*Some common units and their equivalents are listed in Table 1.1. You should be able to use the information, but you will **not** be responsible for memorizing the table. The Table will be given to you during quizzes and exams.*

**Table 1.1 Some Common Units and Their Equivalents**

Length	$1 \text{ m} = 100 \text{ cm}$	$1 \text{ m} = 1000 \text{ mm}$	$1 \text{ cm} = 10 \text{ mm}$	$1 \text{ km} = 1000 \text{ m}$	$1 \text{ nm} = 10^{-9} \text{ m}$
	$1 \text{ \AA} = 10^{-10} \text{ m}$	$1 \text{ in} = 2.54 \text{ cm}$	$1 \text{ ft} = 30.48 \text{ cm}$	$1 \text{ mi} = 1.61 \text{ km}$	$1 \text{ yd} = 0.91 \text{ m}$
	$1 \text{ ft} = 12 \text{ in.}$				
Mass	$1 \text{ kg} = 1000 \text{ g}$	$1 \text{ g} = 1000 \text{ mg}$	$1 \text{ lb} = 454 \text{ g}$	$1 \text{ kg} = 2.20 \text{ lb}$	$1 \text{ oz} = 28.35 \text{ g}$
	$1 \text{ lb} = 16 \text{ oz}$				
Volume	$1 \text{ L} = 1000 \text{ mL}$	$1 \text{ mL} = 1 \text{ cm}^3$	$1 \text{ qt} = 0.946 \text{ L}$	$1 \text{ gal} = 3.78 \text{ L}$	
Energy	$1 \text{ cal} = 4.18 \text{ J}$				
Temperature	$^{\circ}\text{F} = 1.8^{\circ}\text{C} + 32$	$^{\circ}\text{C} = (^{\circ}\text{F} - 32)/1.8$	$\text{K} = ^{\circ}\text{C} + 273.15$		

### Worked Example 1-3

Write conversion factors for each of the following equalities or statements:

- a) 1 g = 1000 mg                      b) 1 foot = 12 inches  
c) 1 quart = 0.946 liter              d) The accepted toxic dose of mercury is 0.30 mg per day.

#### Solution

Equality	Conversion factor	Conversion factor
1 g = 1000 mg	$\frac{1 \text{ g}}{1000 \text{ mg}}$	$\frac{1000 \text{ mg}}{1 \text{ g}}$
1 foot = 12 inches	$\frac{1 \text{ ft.}}{12 \text{ in.}}$	$\frac{12 \text{ in.}}{1 \text{ ft.}}$
1 quart = 0.946 liter	$\frac{1 \text{ qt.}}{0.946 \text{ L}}$	$\frac{0.946 \text{ L}}{1 \text{ qt.}}$
The accepted toxic dose of mercury is 0.30 mg per day.	$\frac{0.30 \text{ mg}}{1 \text{ day}}$	$\frac{1 \text{ day}}{0.30 \text{ mg}}$

## 1.7 | Problem Solving in Chemistry - Dimensional Analysis

**Dimensional analysis** is a general method for solving numerical problems in chemistry. In this method we follow the rule that when multiplying or dividing numbers, we must also multiply or divide units.

Solving problems by dimensional analysis is a three-step process.

1. Write down the given measurement; number *with* units.
2. Multiply the measurement by one or more conversion factors. The unit in each denominator must cancel (match) the preceding unit in each numerator.
3. Perform the calculation and report the answer to the proper significant figures based on numbers given in the question (data), not conversion factors used.

**Worked Example 1-4**

Convert 0.455 km to meters.

**Solution**

To convert kilometers to meters, we could use the following equality:

$$1 \text{ km} = 1000 \text{ m} \text{ (See Table 1.1)}$$

The corresponding conversion factors would be:

$$\frac{1 \text{ km}}{1000 \text{ m}} \quad \text{and} \quad \frac{1000 \text{ m}}{1 \text{ km}}$$

We select the conversion factor to cancel kilometers, leaving units of meters.

$$0.455 \cancel{\text{ km}} \times \frac{1000 \text{ m}}{1 \cancel{\text{ km}}} = 455 \text{ m}$$

The number of significant figures in your answer reflect 0.455 km. The exact conversion factor does not limit the number of significant figures in your answer.

**Worked Example 1-5**

Convert 4.5 weeks to minutes.

**Solution**

$$4.5 \cancel{\text{ wk}} \times \frac{7 \cancel{\text{ d}}}{1 \cancel{\text{ wk}}} \times \frac{24 \cancel{\text{ h}}}{1 \cancel{\text{ d}}} \times \frac{60 \text{ min}}{1 \cancel{\text{ h}}} = 45000 \text{ min}$$

(45360 rounded to 2 sig figs.)

**Worked Example 1-6**

Convert 2.7 g/mL to lb/L.

**Solution**

We need two conversion factors: one to convert g to lb and the other to convert mL to L. We know that 1 lb = 454 g and 1 L = 1000 mL (See Table 1.1)

$$\frac{2.7 \cancel{\text{ g}}}{1.0 \cancel{\text{ mL}}} \times \frac{1 \text{ lb}}{454 \cancel{\text{ g}}} \times \frac{1000 \cancel{\text{ mL}}}{1 \text{ L}} = 5.9 \text{ lb/L}$$

Remember that the number of significant figures in your answer reflect 2.7. The conversion factors do not limit the number of significant figures in your answer.

### Practice 1-7

Perform each of the following conversions:

- a) Convert 14.7 lb to ounces.                      b) Convert 19.8 lb to kilograms.  
c) Convert 23 m/sec to mi/hr.

**Answer**

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## 1.8 | Density and Specific Gravity

**Density** is the ratio of the mass of a substance to the volume occupied by that substance.

$$\text{density} = \frac{\text{mass of substance}}{\text{volume of substance}} \quad \text{or} \quad d = \frac{m}{V}$$

Density is expressed in different units depending on the phase (form) of the substance. Solids are usually expressed in grams per cubic centimeter ( $\text{g/cm}^3$ ), while liquids are commonly grams per milliliter ( $\text{g/mL}$ ). The density of gases is usually expressed as grams per liter ( $\text{g/L}$ ).

**Worked Example 1-7**

If 10.4 mL of a liquid has a mass of 9.142 g, what is its density?

**Solution**

$$d = \frac{m}{V} \qquad d = \frac{9.142 \text{ g}}{10.4 \text{ mL}} = 0.879 \text{ g/mL}$$

Density can be used as a conversion factor that relates mass and volume, note the different units in the numerator and denominator. Densities can be used to calculate mass if volume is given or calculate volume given mass. For example, we can write two conversion factors for a given density of 1.05 g/mL:

$$\frac{1.05 \text{ g}}{1.00 \text{ mL}} \quad \text{or} \quad \frac{1.00 \text{ mL}}{1.05 \text{ g}}$$

**Worked Example 1-8**

The density of a saline solution is 1.05 g/mL. Calculate the mass of a 377.0 mL sample.

**Solution**

$$d = \frac{m}{V} \qquad m = 377.0 \text{ mL} \times \frac{1.05 \text{ g}}{1.00 \text{ mL}} = 396 \text{ g}$$

**Practice 1-8**

The density of rubbing alcohol is 0.786 g/mL. What volume of rubbing alcohol would you use if you needed 32.0 g?

**Answer**

**Specific Gravity** is the ratio of the density of liquid to the density of water at 4°C, which is 1.00 g/mL. Since specific gravity is a ratio of two densities, the units cancel.

$$\text{specific gravity} = \frac{\text{density of sample (g/mL)}}{\text{density of water (g/mL)}} \quad (\text{No units})$$

An instrument called a *hydrometer* is used to measure the specific gravity of liquids.

### **Worked Example 1-9**

What is the specific gravity of jet fuel if the density is 0.775 g/mL?

#### **Solution**

$$\text{specific gravity} = \frac{0.775 \text{ g/mL}}{1.00 \text{ g/mL}} = 0.775$$

### **Practice 1-9**

A 50.0 mL sample of blood has a mass of 53.2 g.

- Calculate the density of the blood.
- Calculate the specific gravity of the blood.

#### **Answer**

## 1.9 | Temperature Scales

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Temperature, reported in **Fahrenheit** ( $^{\circ}\text{F}$ ) or **Celsius** ( $^{\circ}\text{C}$ ), is used to indicate how hot or cold an object is. The SI unit for reporting temperature is **Kelvin** (**K**).

See the comparison of the three scales:

	<b>Freezing point of water</b>	<b>Boiling point of water</b>	<b>Normal body temperature</b>
Fahrenheit	32 $^{\circ}\text{F}$	212 $^{\circ}\text{F}$	98.6 $^{\circ}\text{F}$
Celsius	0 $^{\circ}\text{C}$	100 $^{\circ}\text{C}$	37 $^{\circ}\text{C}$
Kelvin	273 K	373 K	310 K

The following formulas show the conversions:

$$\text{Fahrenheit to Celsius: } ^{\circ}\text{C} = \frac{(^{\circ}\text{F} - 32)}{1.8}$$

$$\text{Celsius to Fahrenheit: } ^{\circ}\text{F} = 1.8 ^{\circ}\text{C} + 32$$

$$\text{Celsius to Kelvin: } \text{K} = ^{\circ}\text{C} + 273$$

### Practice 1-10

Complete the following table.

Fahrenheit	Celsius	Kelvin
88 $^{\circ}\text{F}$		
	-55 $^{\circ}\text{C}$	
		469K

**Answer**

## 1.10 | Heat and Specific Heat

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Heat and temperature are both a measure of energy. Heat, however, is not the same as temperature. **Heat** measures the *total energy*, whereas **temperature** measures the *average energy*. A gallon of hot water at 200°F has much more heat energy than a teaspoon of hot water at same temperature.

Heat can be measured in various units. The most commonly used unit is calorie (cal). The **calorie** is defined as the amount of heat required to raise the temperature of 1 gram of water by 1°C. This is a small unit, and more often we use kilocalories (kcal).

$$1 \text{ kcal} = 1000 \text{ cal}$$

Nutritionist use the word “Calorie” (with a capital “C”) to mean the same thing as kilocalorie.

$$1 \text{ Cal} = 1000 \text{ cal} = 1 \text{ kcal}$$

The unit of energy in SI unit is **joule** (pronounced “jool”), which is about four times as big as the calorie:

$$1 \text{ cal} = 4.184 \text{ J}$$

### Specific Heat

Substances change temperature when heated, but not all substances have their temperature raised to the same extent when equal amounts of heat are added.

**Specific Heat** is the amount of heat required to raise the temperature of one gram of a substance by one degree Celsius. It is measured in units of cal/g·°C or J/g·°C.

(Recall; 1 cal is required to raise the temperature of 1 gram of water by 1°C, the specific heat of water is therefore: 1.00 cal/g·°C, or 4.184 J/g·°C).

Specific heats for some substances in various states are listed in the following table. A substance with a high specific heat is capable of absorbing more heat with a small temperature change than a substance with lower specific heat.

## Specific Heats for Some Common Substances

Substance	Specific Heat (J/g·°C)
<b>Solids</b>	
gold	0.128
copper	0.385
aluminum	0.903
ice	2.06
<b>Liquids</b>	
mercury	0.138
methanol	1.77
ethanol	2.42
water	4.18
<b>Gases</b>	
argon	0.518
oxygen	0.915
nitrogen	1.041
steam	2.03

We can calculate the amount of heat gained or lost by a substance using its specific heat, its measured mass, and the temperature change.

$$\text{Amount of heat} = \text{mass} \times \text{specific heat} \times \text{change in temperature}^*$$
$$q = m \times SH \times (T_{\text{final}} - T_{\text{initial}})$$

\* The temperature change could also be written as  $\Delta$  (*delta* T).

If any three of the four quantities in the equation are known, the fourth quantity can be calculated.

### Worked Example 1-10

Determine the amount of heat that is required to raise the temperature of 7.400 g of water from 29.0°C to 46.0°C. The specific heat of water is 4.18 J/g·°C.

#### Solution

$$q = m \times SH \times \Delta T$$
$$q = 7.400 \text{ g} \times 4.18 \text{ J/g}\cdot\text{°C} \times 17.0\text{°C} = 526 \text{ J}$$

**Practice 1-11**

What mass of lead is needed to absorb 348 J of heat if the temp of the sample rises from 35.2°C to 78.0°C? The specific heat of lead is 0.129 J/g·°C.

**Answer**

**Practice 1-12**

It takes 87.6 J of heat to raise the temp of 51.0 g of a metal by 3.9°C. Calculate the specific heat of the metal.

**Answer**

**Practice 1-13**

$4.00 \times 10^3$  J of energy is transferred to 56.0 g of water at  $19^\circ\text{C}$ . Calculate the final temperature of water.  $SH = 4.18 \text{ J/g}\cdot^\circ\text{C}$ .

**Answer**

## Homework Problems

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1.1 Complete the following table.

Decimal notation	Scientific notation	Number of significant figures
400,000		
0.000600		
21,995,000		
0.05050		
	$7.28 \times 10^3$	
	$3.608 \times 10^{-5}$	
	$9.4090 \times 10^4$	
	$1.5 \times 10^{-3}$	

1.2 Perform the following calculations to correct number of significant figures.

- $4.6 \times 0.00300 \times 193$
- $8.88 \div 99.40$
- $(7.120 \times 10^{-3}) \div (6.000 \times 10^{-5})$
- $(5.92 \times 10^3) \times 3.87 \div 100$

1.3 Perform the following calculations to correct number of significant figures.

- $102 - 5.31 - 0.480$
- $(3.42 \times 10^{-4}) + (5.007 \times 10^{-4})$
- $7.8 - (8.3 \times 10^{-2})$
- $(3.8 \times 10^6) - (8.99 \times 10^6)$

1.4 Perform the following conversions. Show your set ups.

- 683 nanometer (nm) to angstrom (Å)
- 520 mi/h to m/sec
- $0.714 \text{ g/cm}^3$  to  $\text{lb/ft}^3$
- $-164^\circ\text{C}$  to  $^\circ\text{F}$

1.5 A physician has ordered 37.5 mg of a particular drug over 15 minutes. If the drug was available as 2.5 mg/mL of solution, how many mL would you need to give every 15 seconds?

- 1.6 What is the density of a metal sample if a 15.12-g sample is added into a graduated cylinder increased the liquid level from 35.00 mL to 40.60 mL?
- 1.7 The density of copper is  $8.96 \text{ g/cm}^3$ . You have three different solid samples of copper. One is **rectangular** with dimensions 2.3 cm x 3.1 cm x 8.0 cm. The second is a **cube** with edges of 3.8 cm. The third is a **cylinder** with a radius of 1.5 cm and a height of 8.4 cm. Calculate the mass of each sample.
- 1.8 A 50.00-g sample of metal at  $78.0^\circ\text{C}$  is dropped into cold water. If the metal sample cools to  $17.0^\circ\text{C}$  and the specific heat of metal is  $0.108 \text{ cal/g}\cdot^\circ\text{C}$ , how much heat is released?