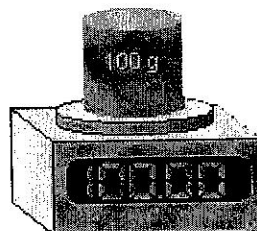


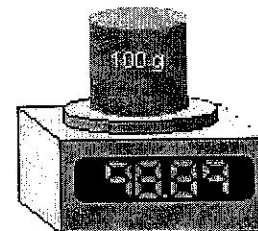
# Accuracy and Precision

**Accuracy** indicates how close a measurement is to the accepted value. For example, we'd expect a balance to read 100 grams if we placed a standard 100 g weight on the balance. If it does not, then the balance is inaccurate.

Accurate



Inaccurate



**Precision** indicates how close together or how repeatable the results are. A precise measuring instrument will give very nearly the same result each time it is used.

There are several ways to report the precision of results. The simplest is the range (the difference between the highest and lowest results) often reported as a  $\pm$  deviation from the average.

A better way, but one that requires statistical analysis would be to report the standard deviation.



More precise	
Trial #	Mass (g)
1	100.00
2	100.01
3	99.99
4	99.99
<b>Average</b>	<b>100.00</b>
<b>Range</b>	<b><math>\pm 0.01</math></b>
<b>Std. Dev.</b>	<b>0.05</b>

Less precise	
Trial #	Mass (g)
1	100.10
2	100.00
3	99.88
4	100.02
<b>Average</b>	<b>100.00</b>
<b>Range</b>	<b><math>\pm 0.11</math></b>
<b>Std. Dev.</b>	<b>0.09</b>

Which of the following sets of data is more precise, based on its range?

<input type="radio"/> Set A	<input type="radio"/> Set B
15.32	32.56
15.37	32.55
15.33	32.48
15.38	32.53
15.35	32.55

Check

Both accuracy and precision affect how many significant digits can be reported. Manufacturers will usually specify the accuracy and precision to be expected from their equipment as a  $\pm$  uncertainty.

It is quite possible for an instrument to be precise, but inaccurate. For example, consider the chains used to measure the first down in a football game. They are supposed to be ten yards long. But what if they were only 9 yards, 35 inches? You would certainly get the same **precise** measurement each time you used the chains, but you wouldn't be getting the correct **accurate** measurement. Both teams would not

have to go quite ten yards to get a first down, and the error is so small you probably wouldn't even notice it. However, there probably have been football games played where one inch would have made a difference to the outcome of the game. In science and in football, our measurements should be both accurate and precise.

Of course, even precise and accurate equipment can be used incorrectly. If the chains were the proper ten yards long, it would still be possible to get an imprecise measurement for first downs. The chains must be stretched tightly, and they must be marked from the proper location on the yard line markers.

The thermometers found in high school labs are often more precise than they are accurate. It is quite easy to read a thermometer to the nearest 0.2 °C. However, the overall calibration can often be off by a degree or more. The temperature shown on the thermometer at the right can be read to 34.0 °C. However, if the thermometer is not of high quality, it would be easy for the real temperature to be off by as much as a degree or more. In other words, the temperature could really be as high as 35 or as low as 33 °C, although the thermometer reads 34.0.



This thermometer reads 34.0 degrees. But, if it is not accurately calibrated, this may not be the true temperature.

In which of the following examples would the precision of a thermometer be more important than its accuracy?

- a) determining the identity of an unknown compound by comparison of its measured melting point to a reference table
- b) measuring the temperature **change** when a chemical is added to water

Check

Try the section quiz to check your understanding of this topic.

---

[Return to MicroScale Index](#)

## MicroScale 25 Temperature and Solubility

[TG Description](#)

[Data Sheet](#)

[Introduction](#)

[Procedure](#)

[Questions](#)

[Safety](#)

[TG Answers](#)

[TG Data Table](#)

[TG Lab Hints](#)

[TG Materials](#)

[TG Reference](#)

---

### TG Description

A systematic study of the variation in the water solubility of a soluble salt with temperature is accomplished by filling the wells of a 12-well strip with solutions saturated at various temperatures and then cooling the strip, to near 0 °C.

[Go to Top](#)

---

### Data Sheet

Make a graph of your own results using the temperature on the x-axis and the height of solid on the y-axis. Draw a smooth line to indicate the pattern formed. Is the line straight or curved?

[Go to Top](#)

---

### Introduction

The ability of a substance to dissolve is called its solubility. The solubility of a substance in water is the mass of the substance that will dissolve in 100 mL of water.

In this experiment you will see how the solubility of a solid changes as the temperature changes. Rather than determine the actual mass of the material dissolved at various temperatures you will use the volume of solid that was dissolved as an indication of the mass of solid that was dissolved.

To find how much solid is dissolved at various temperatures you will allow a hot saturated solution to cool and at specified temperatures remove a small sample with a disposable pipet and fill one well of a 12-well strip. Upon cooling excess solid will appear. The height of this solid in your well will be an indication of how much solid was dissolved at the sampled temperature.

After you complete this Exploration, you will be able to:

1. Find out how much solid is dissolved in a sample of saturated solution at a certain temperature.
2. Prepare and read a solubility graph.

[Go to Top](#)

---

### Procedure

Several baths will be set up in the laboratory (10°C; 25°C; 40°C; 55°C; 70°C; and 85°C). Each will be labeled and will contain a large test tube filled with a saturated solution of  $\text{KClO}_3$ . (The amount of  $\text{KClO}_3$  dissolved in the liquid in each tube is a function of the temperature of the bath. For  $\text{KClO}_3$ , the higher the temperature, the more  $\text{KClO}_3$  is dissolved.)

Bath temperature °C:	10	25	40	55	70	85
Well number	1	2	3	4	5	6

Select a bath. Place a Beral pipet in the tube, and alternately fill and empty it with squeezing several times. Repeat this squeezing over a minute or so to permit the liquid and the pipet to reach thermal equilibrium. Use the table to identify the appropriate well to fill. Fill the pipet about 1/3rd full with solution. Quickly fill the appropriate well of a 12-well strip with the solution. Return unused solution back to the test tube in the bath. Check the temperature of the bath. Record the temperature.

Repeat the procedure for each bath available.

When all the samples have been obtained place the 12-well strip in a shallow pan with a shallow level of ice water. After the strip has been sitting 2 minutes use a toothpick to gently stir each well. Allow the strip to set another 3 to 5 minutes and remove the strip from the cold water. Rapidly measure and record the height of the solid in each well to the nearest 0.5 mm. (Alternately, arbitrarily assign a value from 0 to 10 the the apparent height of solid in the strip with 0 = no solid and 10 = well filled with solid.)

Clean the solid from the strips by placing the strip in warm water, stirring each well with a toothpick. Rinse the wells with distilled water and shake all the water from the wells. Return your 12-well strip to the location specified by your instructor.

[Go to Top](#)

---

### Questions

1. What happens to the solubility of potassium chlorate in water as the temperature is raised?
2. Suppose you heated your sample to 90 °C and all the solid dissolved leaving a clear solution. Would a sample of this clear solution be saturated and thus be suitable for an accurate measure of the solubility at 90 °C? Explain your answer.
3. What indicated that the solution at any given temperature was saturated?
4. Which of the solutions in the 12-well strip would have been saturated after cooling?
5. Most solids that dissolve in water behave like potassium chlorate, although the amount that dissolves