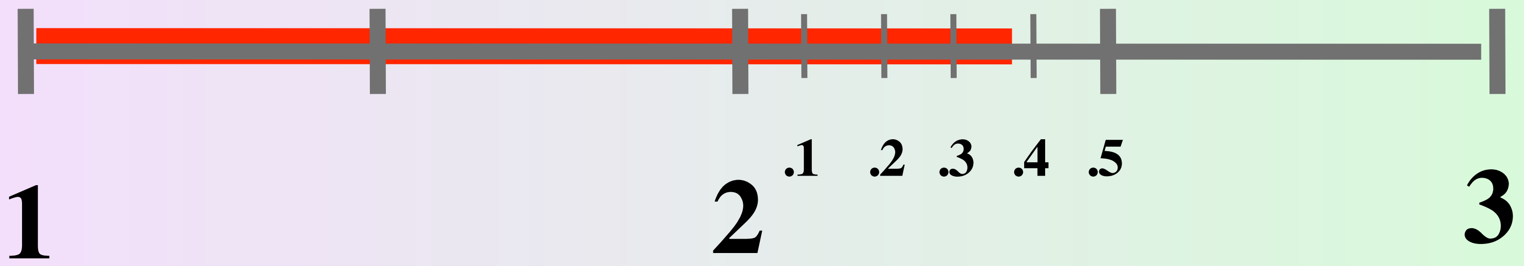
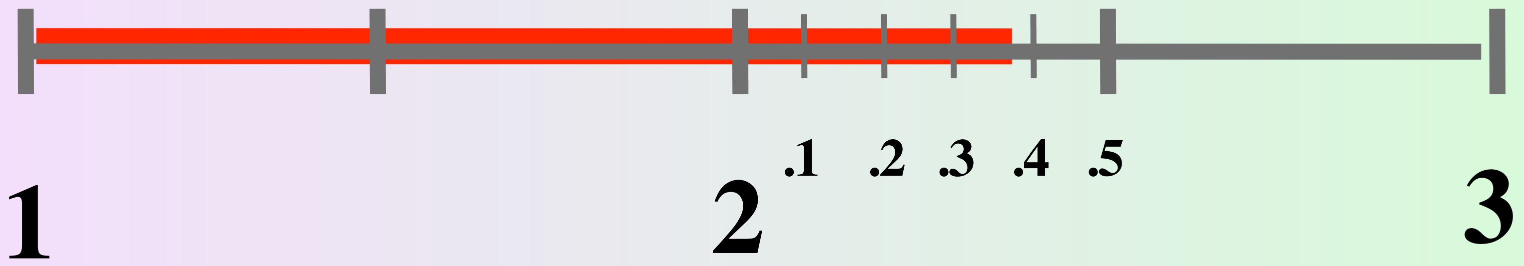
Uncertainty in measurement



2.36mm **2.37mm**

middle value ?

Uncertainty in measurement



2.36mm **2.37mm**

middle value ?

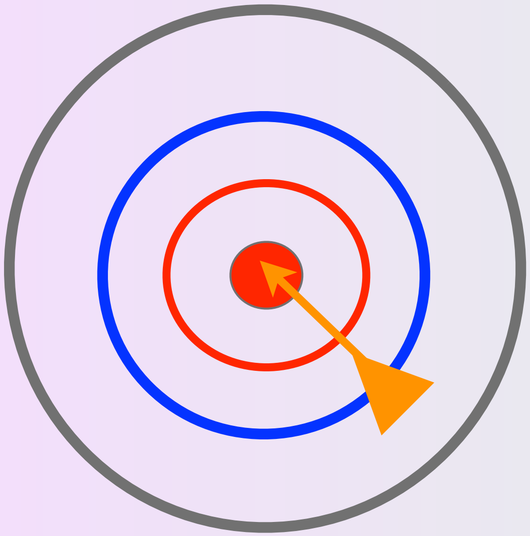
There is uncertainty with this degree of accuracy

Accuracy - the closeness of a measurement to the true value

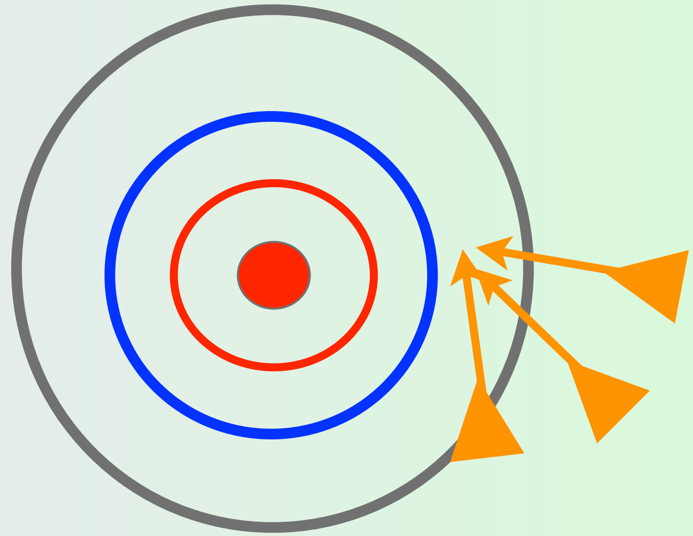
precision - the reproducibility of a series of measurements

A series of measurements can be precise with out being accurate

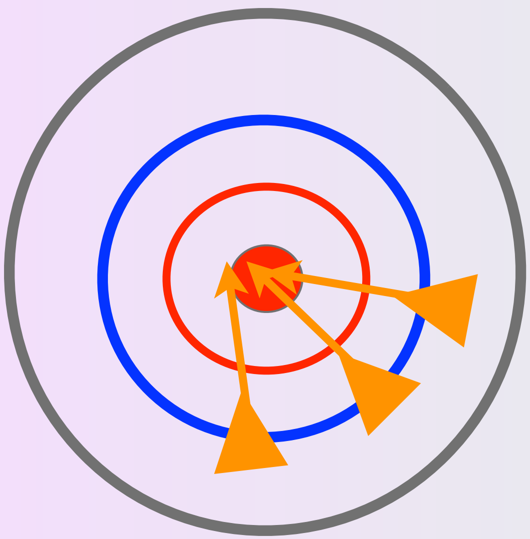
high accuracy



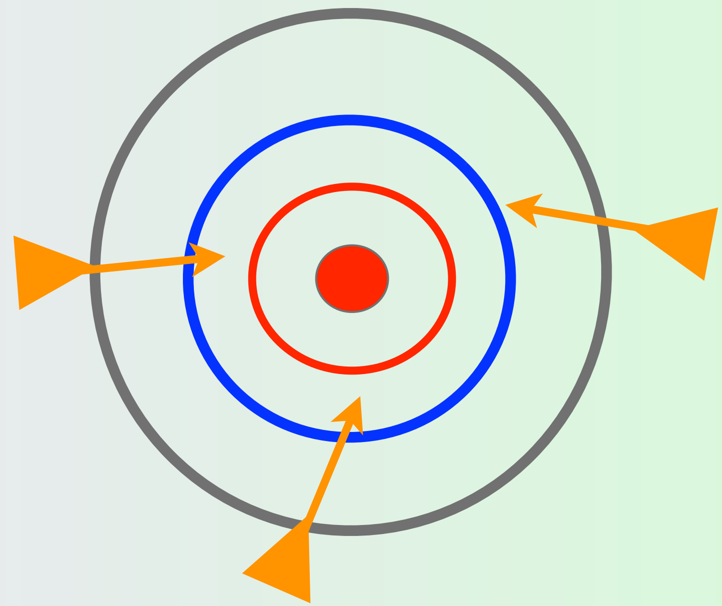
low accuracy / high precision



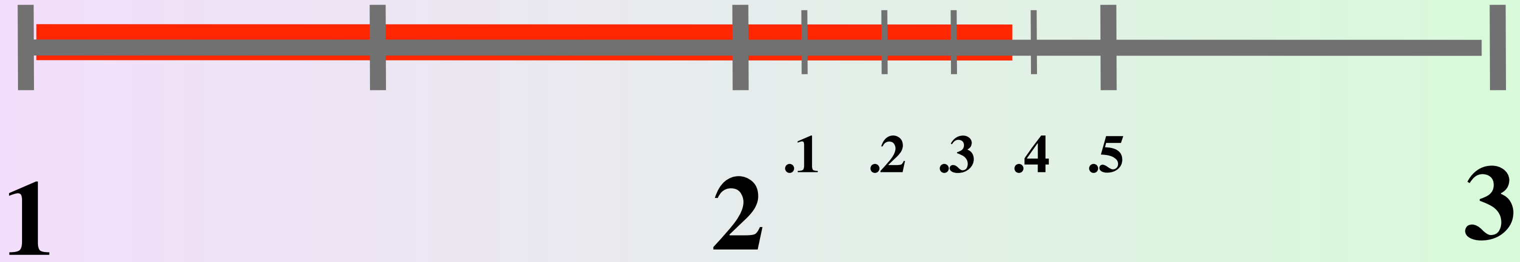
high accuracy/ high precision



low accuracy / low precision



Uncertainty in measurement



2.37mm

The first two measured numbers are called *certain* digits

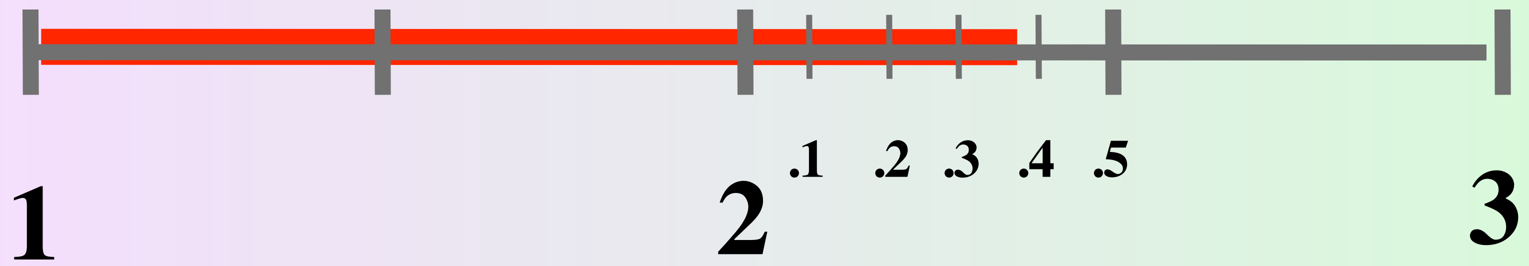
The the digit to the right of the 3 is called an *uncertain* digit

a measurement always has some degree of *uncertainty*

We customarily report a measurement by recording all the certain digits plus the first uncertain digit.

these numbers are called significant figures

Significant Figures



2.37mm

three significant figures in this measurement

Example

What is the difference between the measurements 25.00 ml and 25. ml .

They convey different information:

25.00 ml **a volume between 24.99ml and 25.01 ml**

25. ml **a volume between 24. ml and 26.ml**

Rules for Significant Figures

digits other than zero are always significant

67.8 g	**3 significant figures**
98. g	**2 significant figures**

one or more final zeros used after the decimal point are always significant

4.700 m	**4 significant figures**
82.0 m	**3 significant figures**

Rules for Significant Figures

zeros between two other significant digits are always significant

5.029 cm **4 significant figures**

zeros used solely for spacing the decimal point are not significant

0.00783 ml **3 significant figures**

0.34 g/ml **2 significant figures**

Rules for Significant Figures

If the zeros follow nonzero digits, there is ambiguity if no decimal point is given

300 N **significant figures ?**

300. N **3 significant figures**

300.0 N **4 significant figures**

Avoid ambiguity by expressing measurements in scientific notation

3.0×10^2 N **2 significant figures**

Rules for Significant Figures

Counting numbers and defined constants have an infinite number of significant figures

6 molecules

60 s = 1 min

practice problem 31

Determine the number of significant figures in each measurement.

a. 508.0 L	4 sigs
b. 820400.0 L	7 sigs
c. 1.0200×10^5 kg	5 sigs
d. 807000 kg	3 sigs

Using Significant Figures in Calculations

A result can only be as accurate as the least significant measurement

21.mm - 13.8mm = 7.mm

1 significant figures

Rounding Off Rules

In a series of calculations, carry the extra digits through the final result, then round.

If the digit following the last reportable digit is:

- **4 or less, you drop it** — **1.33 to 1.3**
- **6 or more, you increase the last reportable digit by one** — **1.36 to 1.4**

Rounding Off Rules

If the digit following the last reportable digit is:

• 5, and the next digit is a nonzero round up the last significant figure

round to 3 significant figures

2.5351 to 2.54

Rounding Off Rules

If the digit following the last reportable digit is:

- **5, and the next digit is zero round up the last significant figure**

If the last reportable digit is even, you leave it unchanged

2.5250 to 2.52

If the last reportable digit is odd, you increase it by one.

2.5350 to 2.54

practice problem 33

Round all numbers to four significant figures.

a. 84791 kg	8.479×10^4 kg
b. 38.5432 g	38.54 g
c. 256.75 cm	256.8 cm
d. 4.9356 m	4.936 m

practice problem 34

Round all numbers to four significant figures. Write all answers in scientific notation.

a. 0.00054818 g — 5.482×10^4 g

b. 136758 kg — 1.368×10^3 kg

c. 308659000 mm — 3.087×10^8 mm

d. 2.0145 ml — 2.015×10^0 ml

Adding Significant Figures in Calculations

A result can only be as accurate as the least significant measurement

4.37 g

+ 1.002 g

5.37~~2~~ g **3 significant figures**

Multiplying Significant Figures in Calculations

A result can only be as accurate as the least significant measurement

Volume = l x w x h = (1.87cm)(1.413cm)(1.207cm)

= 3.19cm^3

3 significant figures

Percent Error

Observed value

the value based on laboratory measurements

True value

the value based on accepted references

Absolute error

the difference between the observed value and the true value

(observed value - true value)

Percent Error

$$\% \text{ Error} = \frac{\text{absolute error}}{\text{true value}} \times 100\%$$

Example

the boiling point of methanol is 65˚C. Your measured boiling point of methanol is 66.0˚C.what is the percent error in your measurement.

$$\% \text{ Error} = \frac{\text{absolute error}}{\text{true value}} \times 100\%$$

$$\% \text{ Error} = \frac{66^\circ C - 65^\circ C}{65^\circ C} \times 100\%$$

$$= 1.5\%$$

Graphing

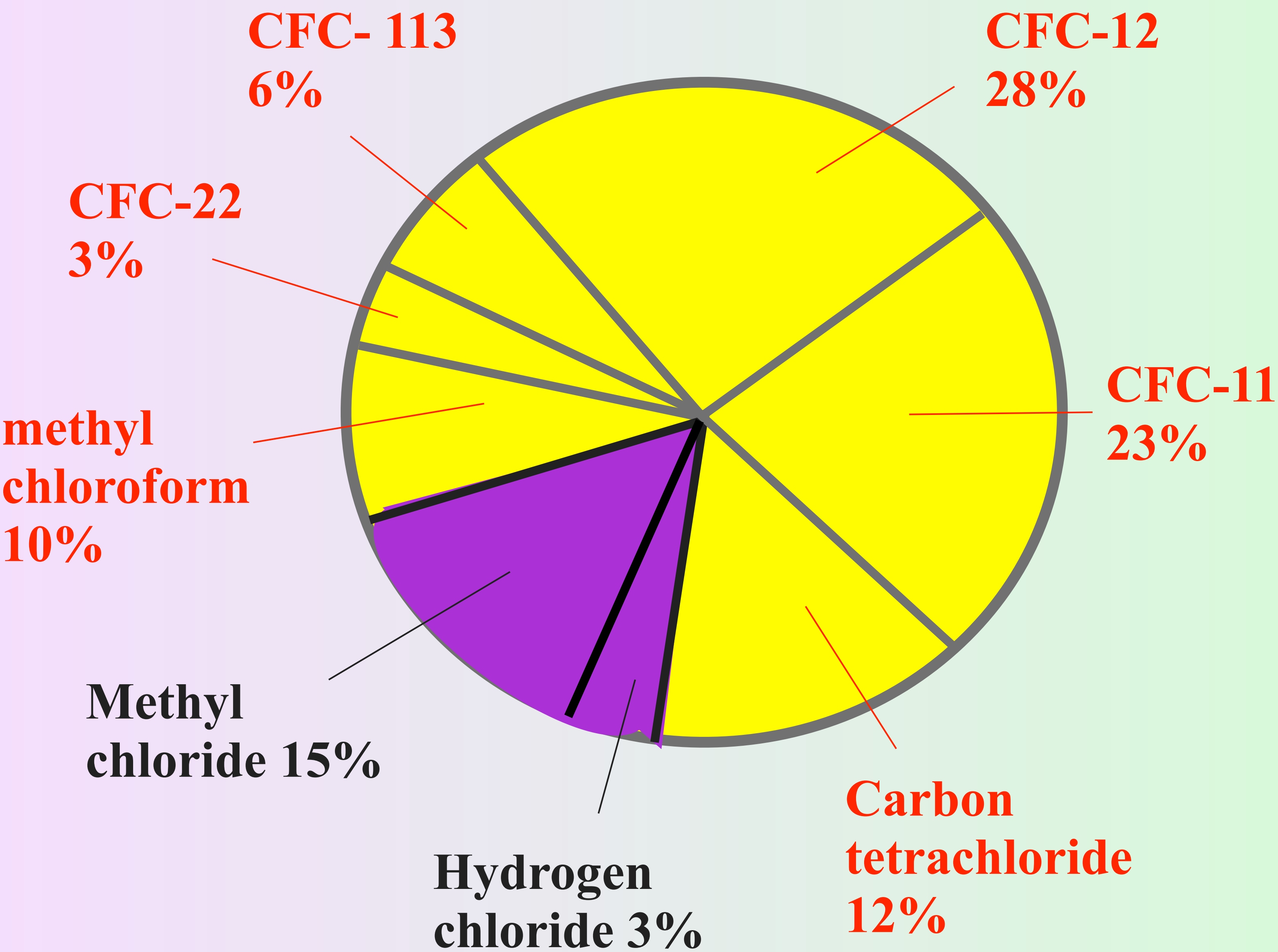
using data to help reveal if a pattern exists

understanding the past , predicting the future

Circle graph

useful for showing parts of a fixed whole

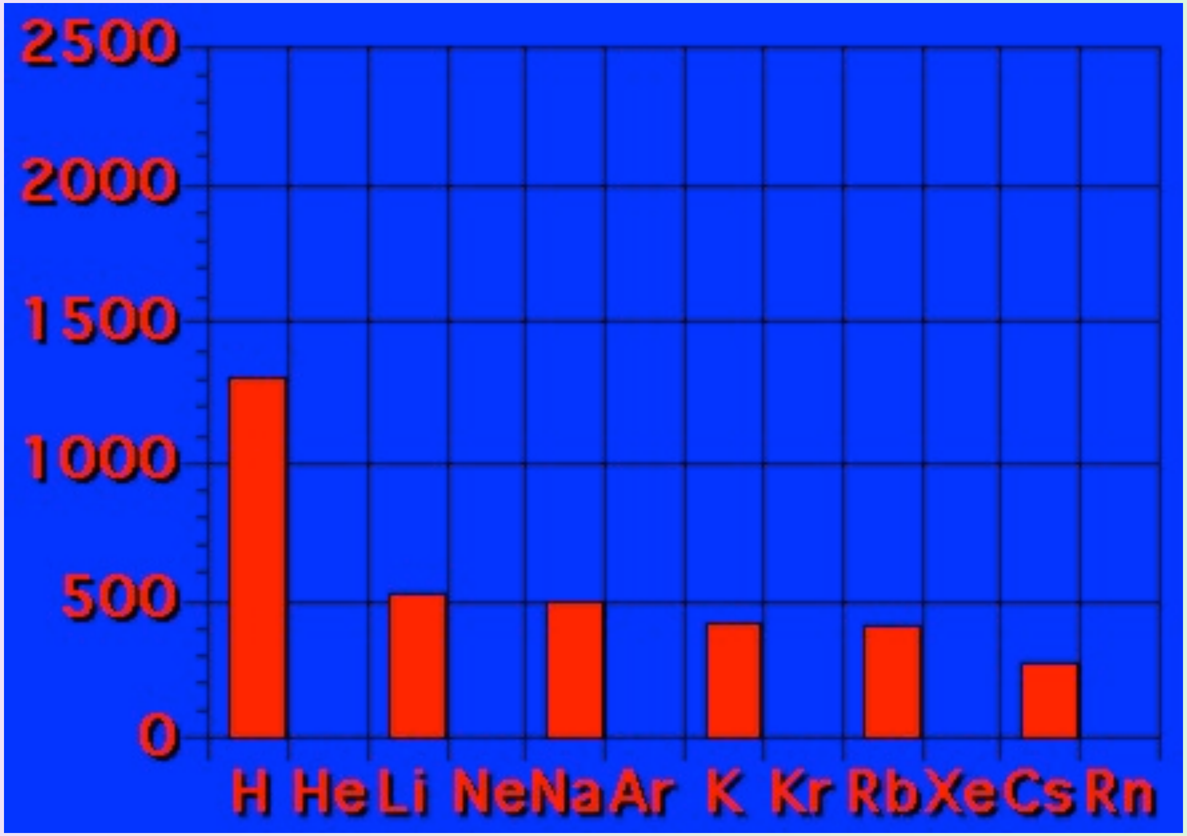
sources of Chlorine in the stratosphere



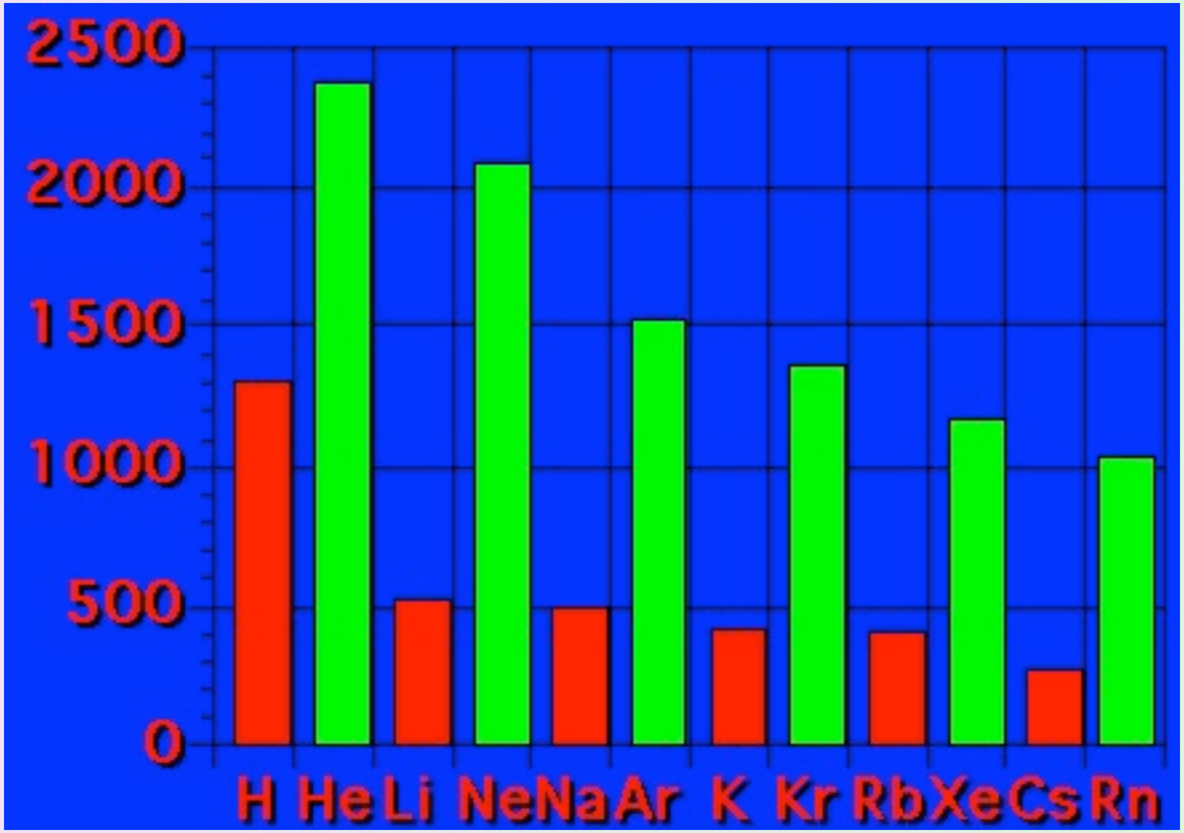
Bar graph

the quantity being measured appears on the vertical axis (*y* - axis). The independent variable appears on the horizontal axis (*x* - axis).

Plot of Ionization Energy

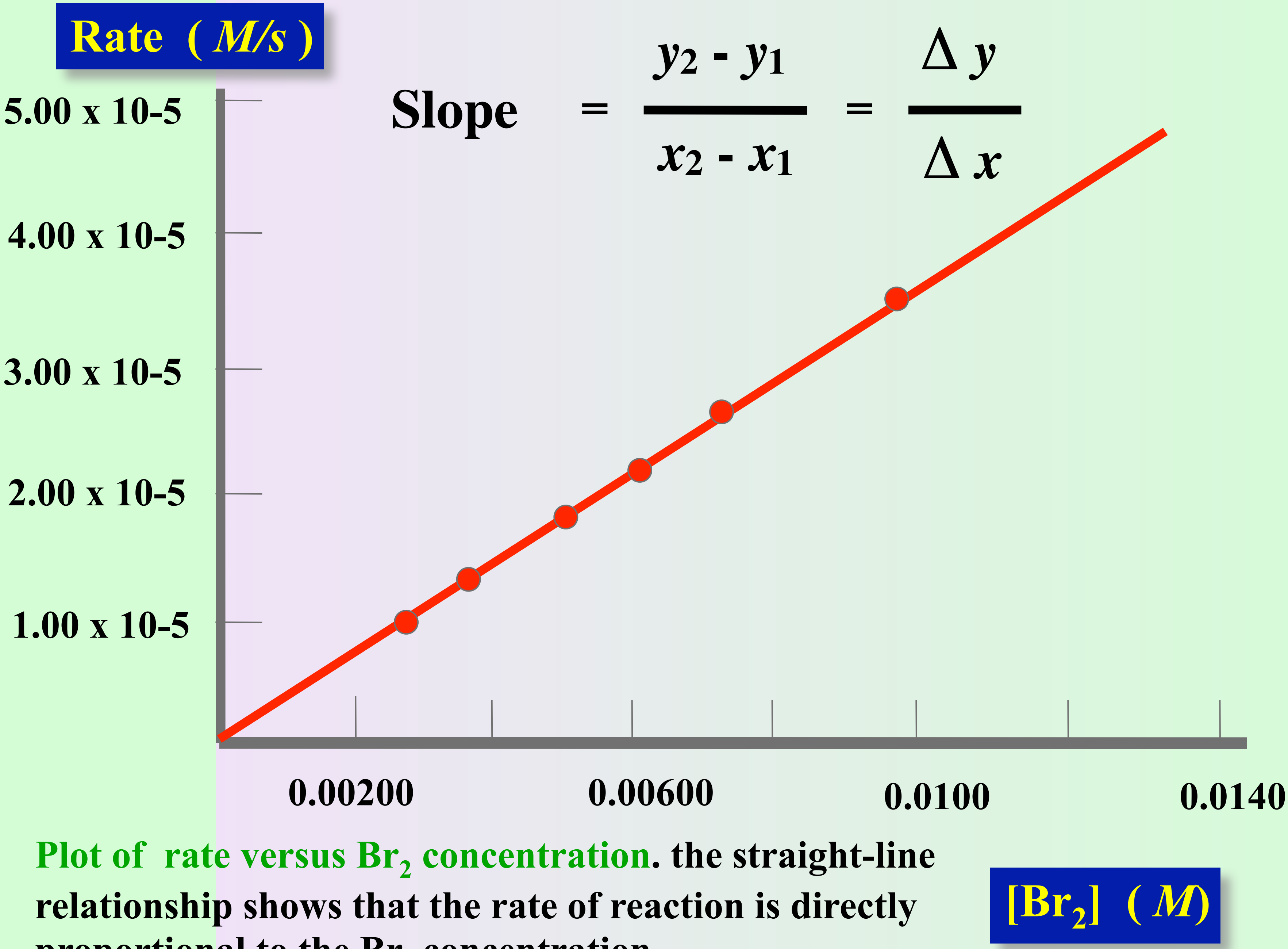


Plot of Ionization Energy



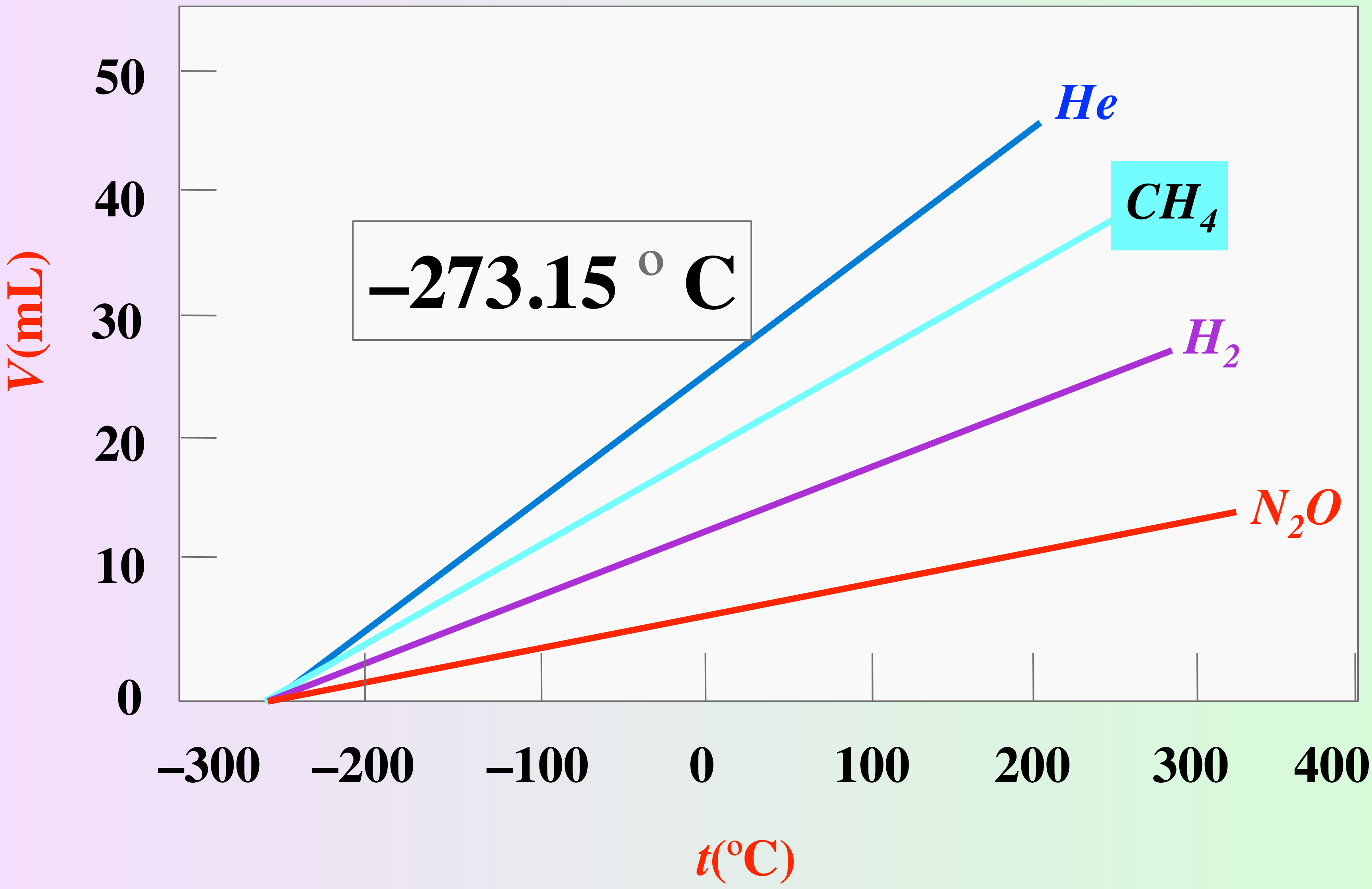
Line graph

The points on a line represent the intersection of two variables. The independent variable appears on the (*x* - axis). The dependent variable appears on the (*y* - axis).



$$\text{Slope} = \frac{y_2 - y_1}{x_2 - x_1} = \frac{\Delta y}{\Delta x}$$

Plot of rate versus Br_2 concentration. the straight-line relationship shows that the rate of reaction is directly proportional to the Br_2 concentration



V(mL)
50
40
30
20
10
0
He
CH4
H2
N2O
–273.15 ° C
–300
–200
–100
0
100
200
300
400
t(°C)