Significant Figures

Rules for Counting Sig. Fig.

1. **All** non-zero digits are significant

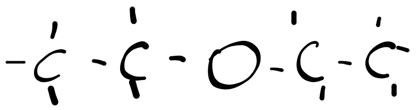
2. Zeroes
a) zeroes between non-zero digits are significant
Ex. 507

b) leading zeroes are not significant
Ex. 0.00004

c) Trailing zeroes to the right of a number are significant **if the number has a decimal point**. If the number ends in zero and has no decimal point, we assume that the trailing zeroes are not significant.
Ex. 480.0 (4 sig figs)
Ex. 4800 (2 sig figs)

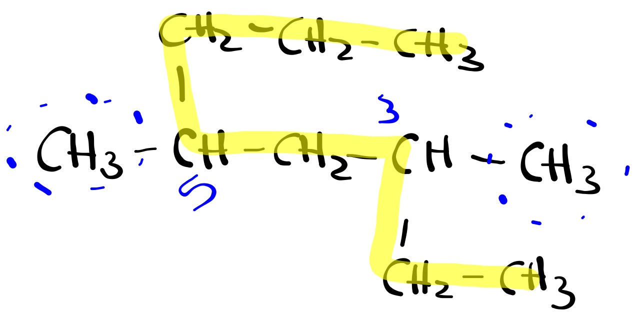
How many significant figures in the following?

a) 38.4703 mL - ___ sig. figs
b) 0.0052 g - ___ sig. figs
c) 0.05700 s - ___ sig. figs
d) 6.19×10^8 years - ___ sig. figs



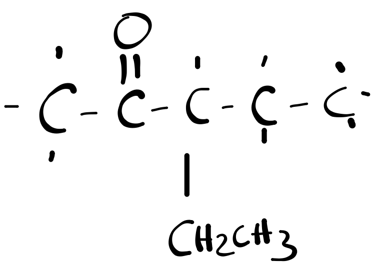
5-phenyl-1-pentene

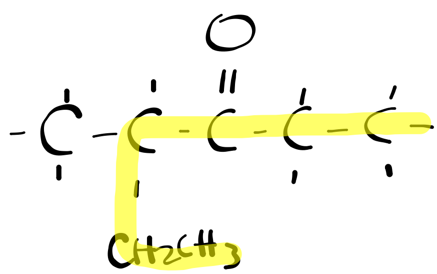
a)



3,5-dimethyloctane

4. a) ethyl pentanone





Significant Figures and Calculations

1. Multiplication and Division

The result of the operation is reported as having as many significant figures as the measurement with the fewest significant figures

Ex. (6.221 cm) x (5.2 cm) =

2. Addition and Subtraction

The result of the operation is reported to the same number of **decimal places** as that of the term with the least number of **decimal places**

Ex.
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    20.4
     1.322
+   83
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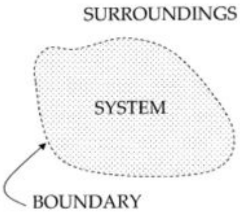
Energy Changes

Energy - is the ability to do work.

A chemical substance or **system** is a group of chemicals being studied, separated from the surroundings by a boundary.

The **surroundings** are defined as the environment around the chemical system.



When a system loses thermal energy, the surroundings will gain that energy.

When a system gains thermal energy, it usually gets it from the surroundings.

Temperature (T) - is a measure of the average kinetic energy of the particles in a system. Units : °C

Kinetic energy - energy of motion

Heat (q) - is a measure of the total kinetic energy of the particles in a system. Units : J (joule)

Heat will always flow from a warm object to a cooler object.

Heat Capacity (C) - is the heat required to change the temperature of a quantity of the substance.

water → 4.19 $\frac{J}{g \cdot °C}$ steam → 2.01 $\frac{J}{g \cdot °C}$

There are two kinds of heat capacity:

1. Specific Heat Capacity (J/g*°C) - is the quantity of heat required to raise the temperature of a unit mass (1 gram) of a substance by a temperature of one degree Celcius.
2. Volumetric Heat Capacity (MJ/m³*°C) - is the quantity of heat required to raise the temperature of a unit volume of a substance by one degree Celcius.

Volumetric heat capacity is used mainly with liquids and gases.

Note: 1 (MJ/(m³*°C) = 1 (kJ/(L*°C))

4.19 MJ/m³ · °C = 4.19 $\frac{kJ}{L \cdot °C}$

The quantity of heat (q) that flows varies directly with the quantity of substance (mass or volume), the specific or volumetric heat capacity (C) and the temperature change(ΔT).

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FORMULA: $q = m\,C\,\Delta T$ or $q = v\,C\,\Delta T$

In calculating q, the heat capacity constant (C) must correspond to the state of matter of the substance.

$$q = mC\Delta T$$

- q: heat (J)
- m: mass (g)
- C: spec. heat capacity $\left(\frac{J}{g\,^\circ C}\right)$
- ΔT: change in temp. (°C)

$$q = vC\Delta T$$

- v: volume (L)
- C: volumetric heat capacity $\left(\frac{kJ}{L\cdot{}^\circ C}\right)$

Example

How much energy does it take to raise the temperature of 50.0 g of copper from 18.0°C to 30.0°C?

$q = ?$

$m = 50.0 g$

$C = 0.385 \frac{J}{g°C}$

$\Delta T = 12.0°C$

$q = mC\Delta T$

$q = (50.0 g)(0.385 \frac{J}{g°C})(12.0°C)$

$q = 231 J$

Determine the amount of energy required to heat 15.00 g of ice from -25.00°C to -8.000°C.

If 1935 J of energy is lost when 100. g of a substance is cooled from 43.0°C to 21.5°C, what is the substance's specific heat capacity?

$q = -1935\,J$
$m = 100.\,g$
$\Delta T = -21.5°C$
$C = ?$

$q = mC\Delta T$

$-1935\,J = (100.\,g)\,C\,(-21.5°C)$

$C = \frac{-1935\,J}{(100\,g)(-21.5°C)}$

$C = 0.900\ \frac{J}{g \cdot °C}$