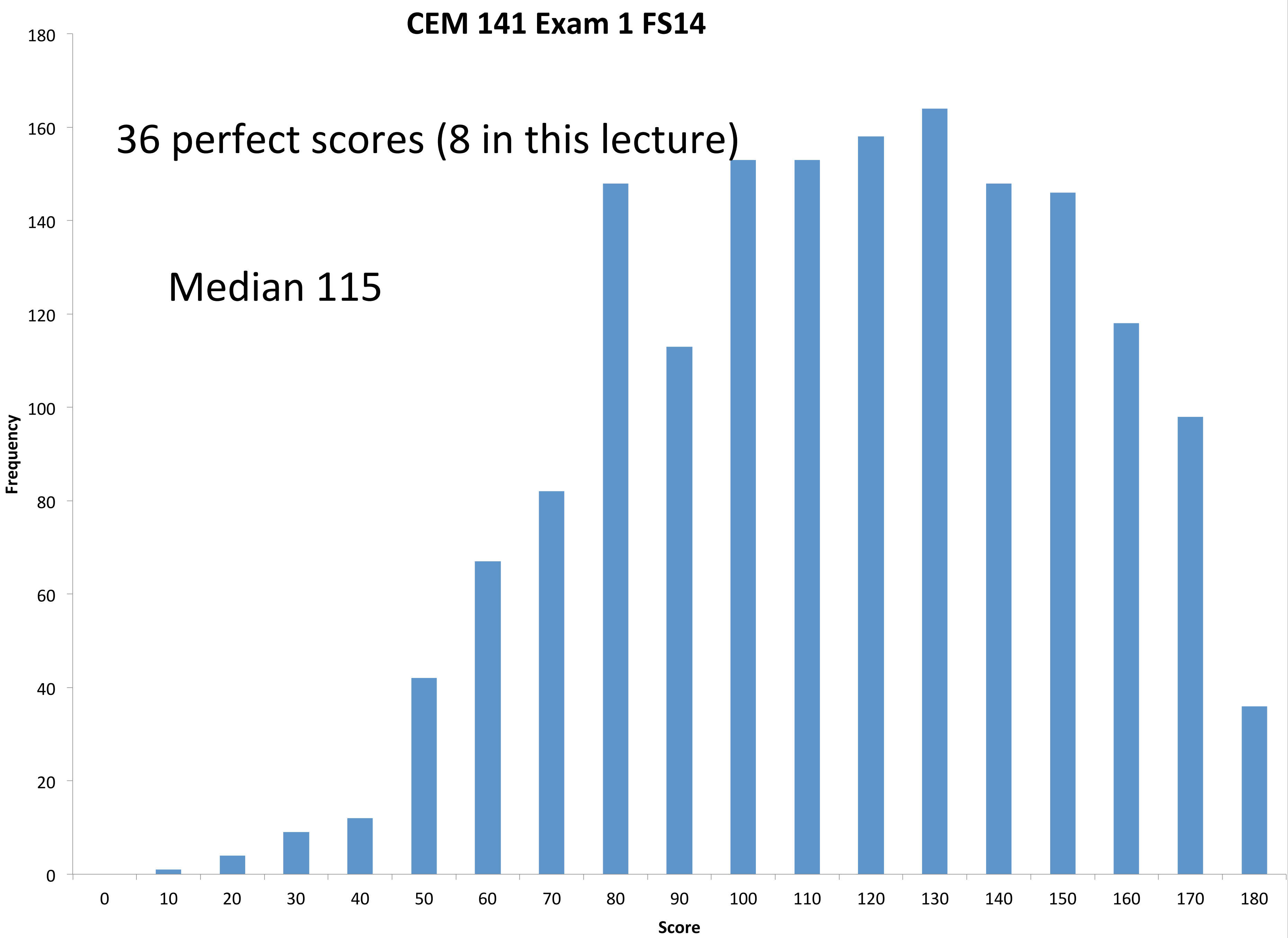

CEM 141 Exam 1 FS14
36 perfect scores (8 in this lecture)
Median 115
Frequency
Score
0
10
20
30
40
50
60
70
80
90
100
110
120
130
140
150
160
170
180

Solution Stoichiometry

Since many reactions occur in solution we often need to know the number of moles of a particular substance in a given volume of solution.

The most common concentration unit is called the molarity and is represented by an upper case M

The molarity of a compound in a solution is the number of moles of that compound per liter of solution

Example 1 solution stoichiometry

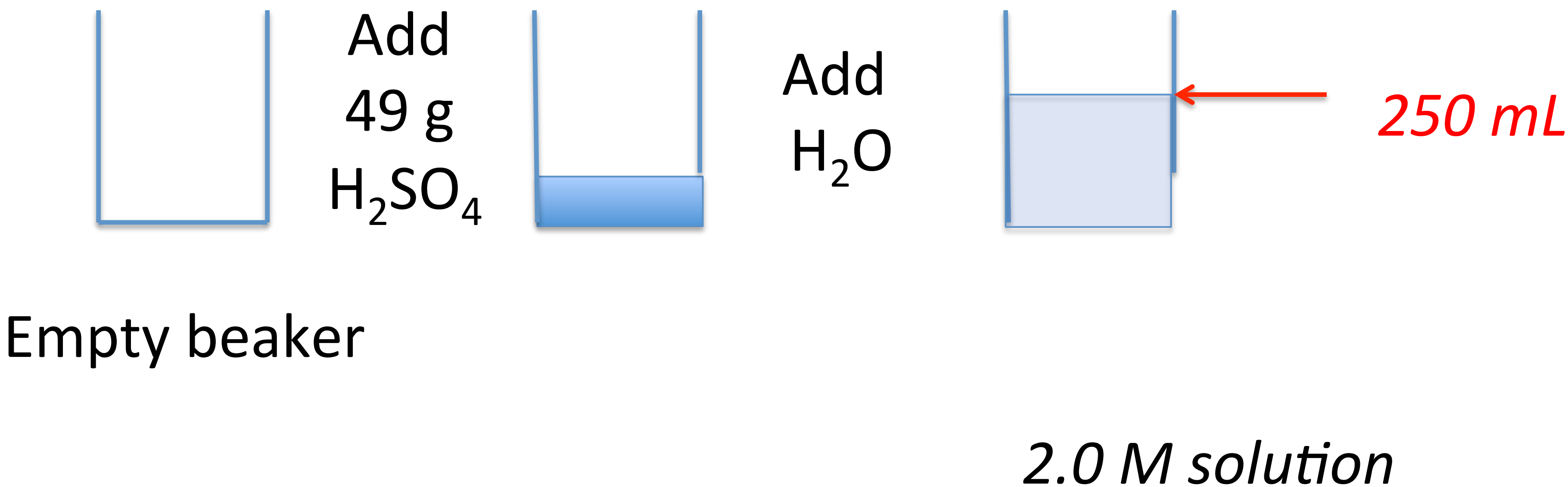
Dissolve 49 g of H_2SO_4 in enough water to make 250 mL of solution.

What is the molarity of the solution?

$$M = \frac{moles\ of\ material}{liters\ of\ solution}$$

Molar mass of H_2SO_4 = 98 g/mol   49 g is ½ mole

$$M = \frac{moles}{L} = \frac{0.50}{0.25} = 2.0$$


Add
49 g
H_2SO_4
Add
H_2O
250 mL
Empty beaker
2.0 M solution

Example 2 solution stoichiometry

What mass of NaOH in grams, is required to make 15.0L of a 0.200 M solution?

$$M = \frac{moles}{liters} = \frac{moles}{V}$$

M = 0.200 *V = 15.0L* $moles = 0.200 \times 15.0 = 3.00$

Molecular mass of NaOH = 40.0 g/mol

$$grams\ NaOH = 3.00 mol \times \frac{40.0 g}{mol} = 120 g$$

Example 3 solution stoichiometry

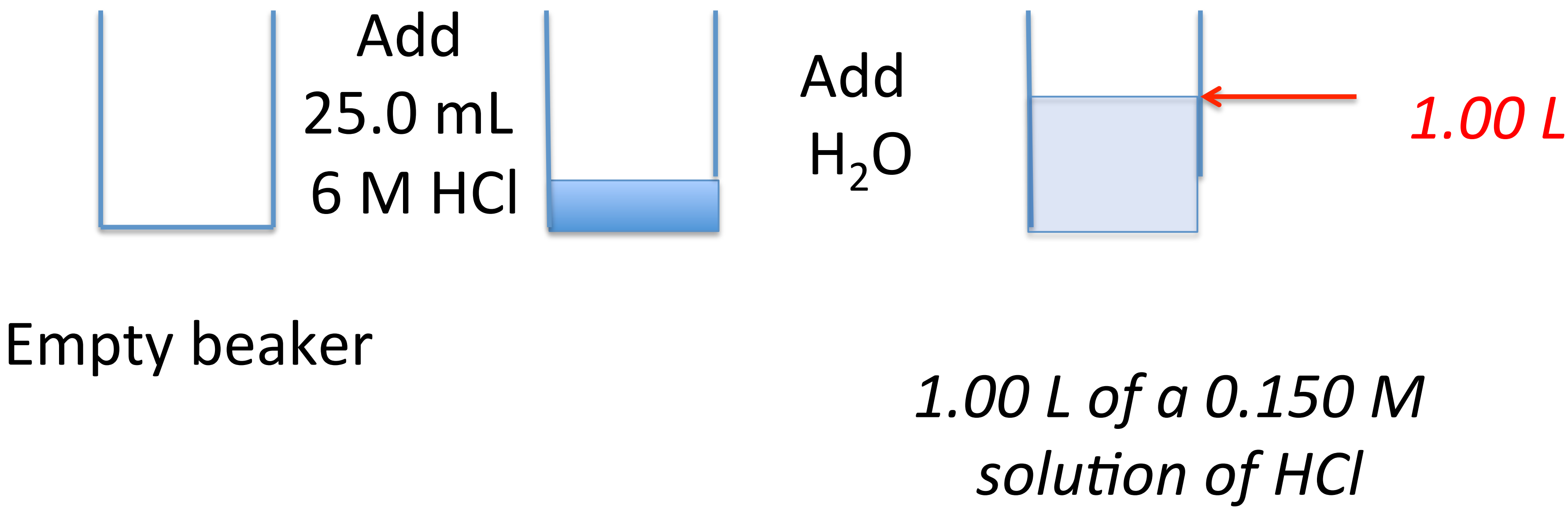
How many mL of a 6.00 M HCl solution does one need to make 1.00 L of a 0.150 M HCl (aqueous) solution?

$$M = \frac{moles}{V}$$

1.00L of a 0.150M solution contains 0.150 moles of HCl

What volume of a 6.00 M HCl solution contains 0.150 moles of HCl?

$$V = \frac{0.150 mol}{6.00 mol / L} = 0.0250 L \equiv 25.0 mL$$


Add
25.0 mL
6 M HCl
Add
H_2O
1.00 L
Empty beaker
1.00 L of a 0.150 M
solution of HCl

Lecture notes 11

Thermodynamics

Developed initially to understand how one can efficiently convert heat into useful work

We will use it to understand the energy changes that take place in a chemical reaction

Thermodynamics

$$\textit{Reactants} \rightarrow \textit{Products}$$

$$\textit{bonds broken} \rightarrow \textit{bonds formed}$$

$$\textit{energy required} \rightarrow \textit{energy released}$$

If the energy required is more than the energy released

endothermic reaction

If the energy released is more than the energy required

exothermic reaction

Some definitions

System: a part of the universe under consideration

Surroundings : everything else

System + Surroundings = Universe

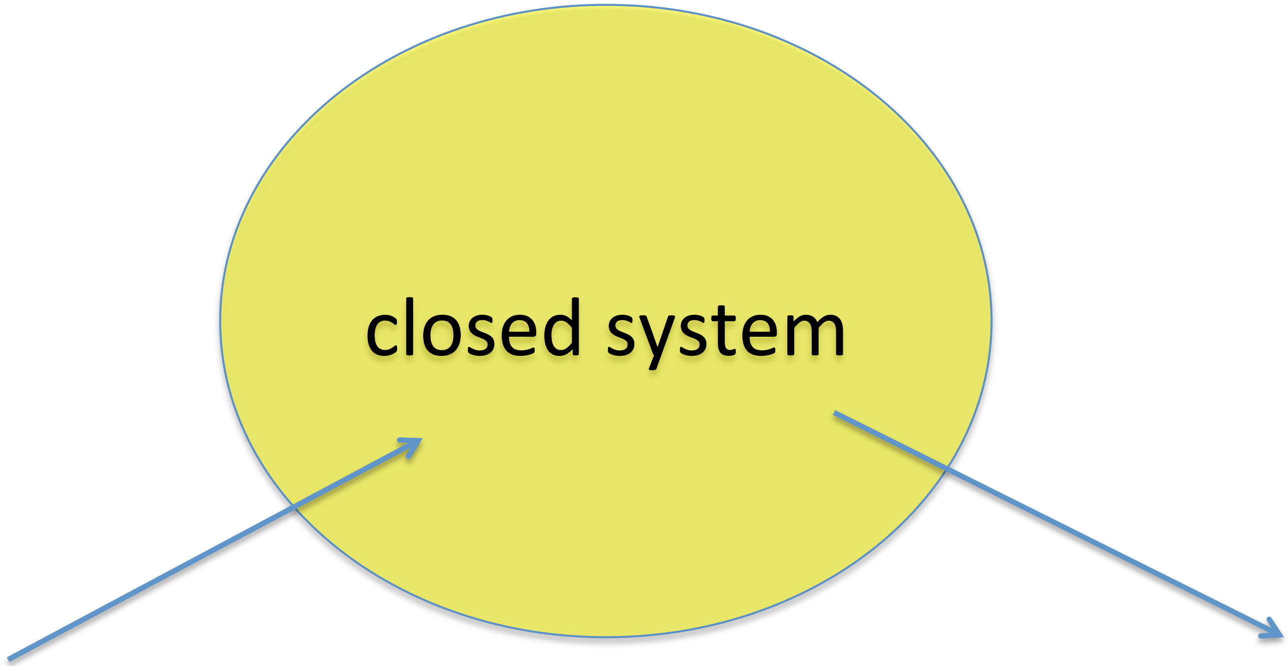
Three types of systems

Open system: can exchange
both
matter and energy with surroundings

Closed system: can exchange
only
energy with surroundings

Isolated system: cannot exchange
energy or matter with the surroundings

$$\Delta E_{system} = heat\ \ transfered + work\ \ done\ \ by/on\ \ system$$



Energy in

Add heat

Do work on system

Energy out

Remove heat

Have the system do work

q measure of heat involved

q

is positive if heat is added to the system

and

negative if heat is removed from the system

W is measure of the work involved

W

is positive if work is done on the system

and

negative if the system does work

First law of Thermodynamics

The energy of the universe is constant

$$\Delta E_{system} + \Delta E_{surroundings} = 0$$

Energy cannot be
created or destroyed
but simply transferred

$$\Delta E_{system} = \Delta E = q + W$$

Suppose we add 40 Joules of thermal energy to a system and the system does 15 Joules of work. What is the change in the energy of the system?

$$\Delta E = q + W$$

$q = +40$ Joules

$W = -15$ Joules

$\Delta E = 40 - 15 = +25$ J

Heat is a manifestation of the
kinetic energy (energy of motion)
of the atoms or molecules
constituting the system

The hotter the material
the
faster the atoms and molecules are moving

This motion is random

Consequently heat is an incoherent form of energy

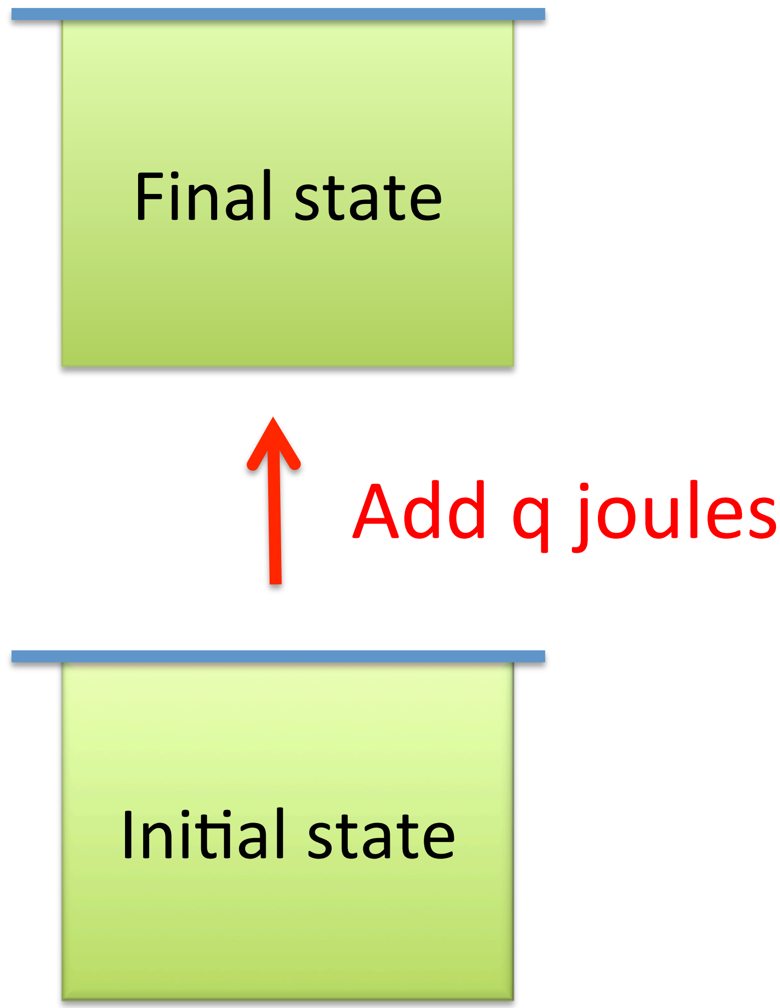
Closed system

Volume doesn't change

Final energy E_f

$E_f - E_i = \Delta E = q$

Initial energy E_i



Closed system

$P_{external}$

Final energy E_f

Volume changes

Add q joules

$E_f - E_i = \Delta E = q + W$

W is negative

$W = -P_{ext}\Delta V$

$P_{external}$

Initial energy E_i

System does work against the external pressure

Energy change when heat is added and the volume of the system changes.

$$work = Force \times distance = \frac{Force}{Area} \times distance \times Area = Presssure \times Volume$$

"pressure volume work" = $-P_{external}\,\Delta V = -P\Delta V$

Volume increases

work is negative (system is doing work)

Volume decreases

work is positive, work is being done on the system

We can measure work by knowing the external pressure and the volume change.

What about heat?

Measure the temperature change in the system and use the specific heat of the system to calculate the heat change.

Calorimeter

What is a specific heat?

The amount of heat required to raise the temperature of 1 gram of material 1 degree Kelvin

Material	specific heat (J/gK)
ice	2.1
water	4.184
steam	2.0
Cu	0.385

$$q(J) = mass(g) \times Temperature\ change(K) \times specific\ heat\left(\frac{J}{gK}\right)$$

Represent specific heat by a lower case “c”

Represent mass of material by a lower case “m”

Represent temperature change by ΔT

$$q(J) = m(g) \times \Delta T(K) \times c\left(\frac{J}{gK}\right)$$

Example 1 lecture notes 11

How much heat is required to raise the temperature of 50 grams of water from 20^0C to 45^0C ?
Specific heat of water = 4.184 $Jg^{-1}K^{-1}$

$$q(J) = m(g) \times \Delta T(K) \times c\left(\frac{J}{gK}\right)$$

$$q = 50.0g \times 25K \times 4.184\frac{J}{gK} = 5230J = 5.23 \times 10^3 J = 5.23kJ$$

Temperature in Kelvin and Celsius are related by

$$T(K)=T(^0C)+273.15$$

$$\Delta T(K)=\Delta T(^0C)$$

Heating curves
(plot of temperature versus heat added)