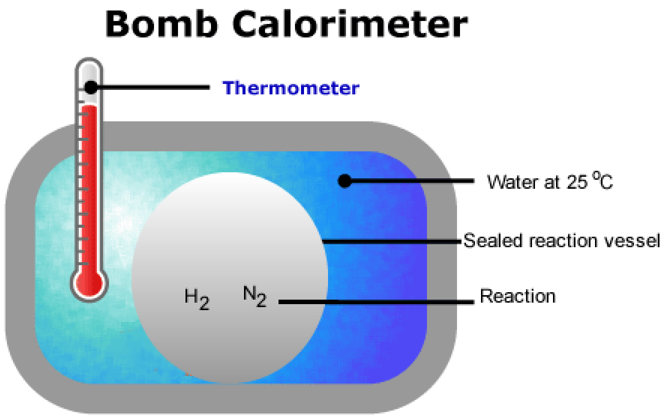
Friday Worksheet **Name: ……………….**

Enthalpy and rate worksheet

1) The Haber process is important for the industrial production of ammonia. The equation for this reaction is given below.

$$N_2(g) + 3H_2(g) \leftrightarrows 2NH_3(g) \qquad \Delta H = -91 \text{ kJ.mol}^{-1}$$

3.90 g of N_2 gas and 0.800 g of H_2 gas react according to the equation above.



The process uses an iron catalyst.

a) Calculate the number of mol of each reactant present.

$n_{nitrogen}$ = 3.90/28.0 = 0.139

$n_{hydrogen}$ = 0.800/2.00 = 0.400

b) Identify the limiting reagent.

Hydrogen gas.

c) Calculate the amount of excess reagent

Since H_2 is the limiting reagent we will use H_2 to calculate the amount of N_2 used.

$\underline{N_2(g) + 3H_2(g) \leftrightarrows 2NH_3(g)}$

According to the stoichiometric ratio for every 3 mol of hydrogen that reacts one mol of nitrogen reacts. So, for 0.400 mol of hydrogen we need 0.400/3 mol of nitrogen.

0.133 mol of nitrogen reacts.

We originally have 0.139 mol and since 0.133 mol reacts an excess of 0.006 mol of nitrogen remains.

d) Will the temperature of the water rise or fall? Explain.

Since ΔH is negative, heat is given out. This heat energy will cause the temperature of the water to rise.

e) What amount of energy, in kJ, is given out by the reaction?

$$N_2(g) + 3H_2(g) \leftrightarrows 2NH_3(g) \qquad \Delta H = -91 \text{ kJ.mol}^{-1}$$

For every 3 mol of hydrogen used, 91 kJ of heat energy is given out.

So the ratio is 91/3 and this ratio remains constant for the equation above. Hence we can write the following expression.

$$\frac{91 \text{ kJ}}{3.00} = \frac{X}{0.400}$$

⇨ x = 91 X 0.4 / 3.00 = 12.1 kJ

f) Explain the impact of using a catalyst under the following headings.
How does a catalyst impact on the:

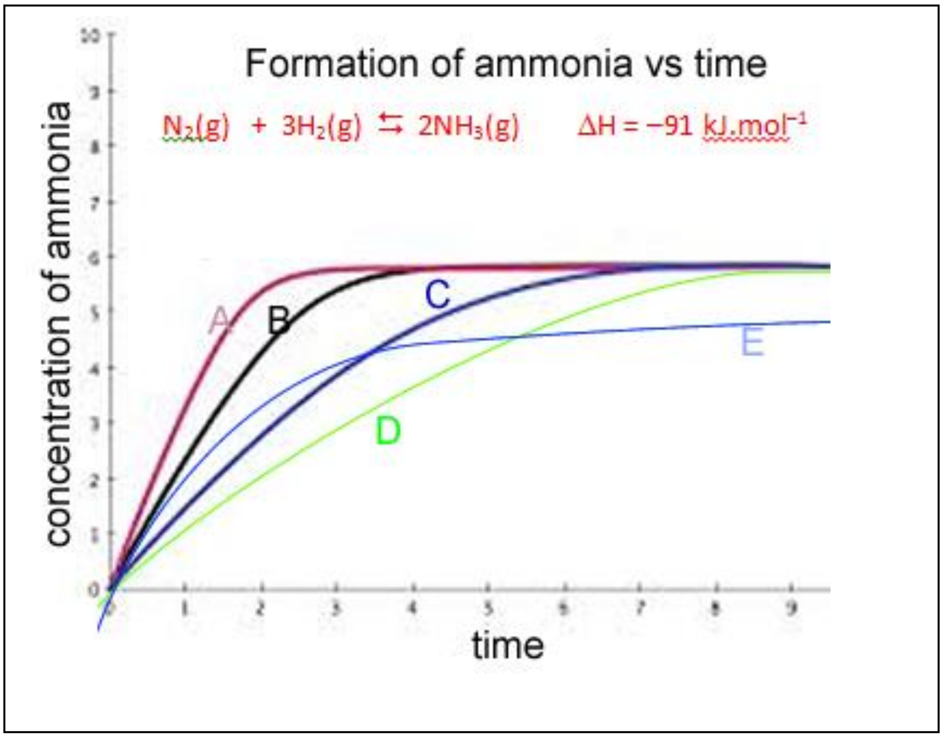
i) number and type of collisions taking place between reactant particles per second,

The number of collisions taking place between reactant particles does not change, however, since the activation energy required is lowered by the catalyst a greater number of these collisions lead to a reaction. The number of fruitful collisions increases.

ii) total amount of heat given out by the reaction,
The total amount of heat given out is not affected by the catalyst.

iii) forward and backward rates of the reaction once equilibrium is achieved,
A catalyst causes a reaction to reach equilibrium quicker. Once at equilibrium the catalyst increases the rate of the forward and backward reactions equally.

g) The reaction is conducted five different times. Each time the reaction is conducted only one variable is modified and the amount of ammonia formed measured against time. Graph C represent the reaction at 25 °C in the absence of an iron catalyst.



Indicate which graph is consistent with the following conditions. Explain your choice

i. The use of a solid iron catalyst.
Graph B. The rate of ammonia production is accelerated reaching the same yield as graph C

ii. The use of a powdered iron catalyst.
Graph A. The powder increases the surface area and so should increase the rate of reaction even more than B.

iii. The reaction taking place at 8 °C in the absence of acatalyst..

Graph D. At lower temperatures the rate of a reaction decreases as particles have less kinetic energy.

iv. The reaction taking place at 500 °C

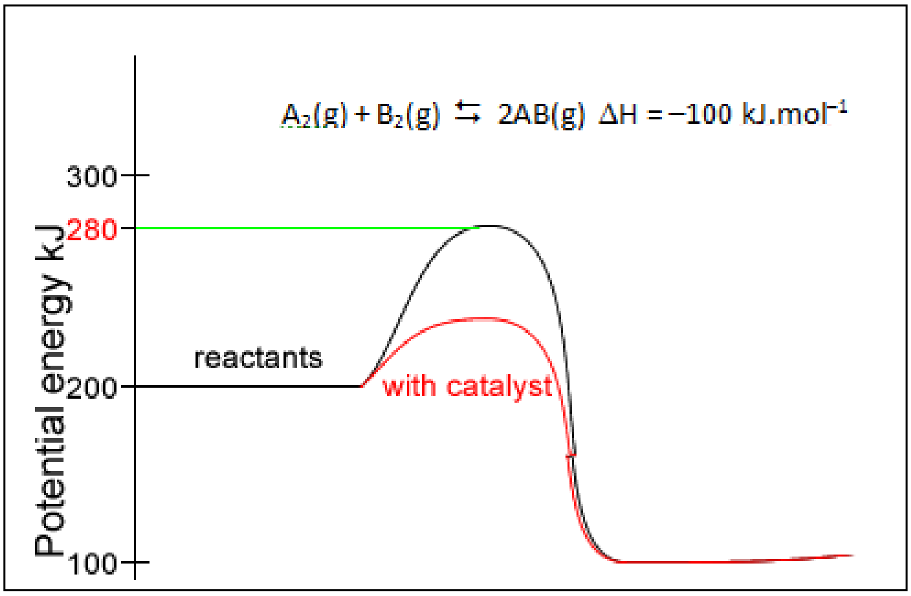
Graph E. An increase in temperature will increase the rate but decrease the yield.

h) One mole of reactant A_2 reacted with one mole of reactant B_2 according to the equation below.

$$A_2(g) + B_2(g) \leftrightarrows 2AB(g) \quad \Delta H = -100\ kJ.mol^{-1}$$

It required 80 kJ of energy to break the bonds of the reactants.

i. Draw an energy profile for the reaction above on the set of axes below.



ii. The same reaction was conducted in the presence of a catalyst. Draw the energy profile on the same set of axes above.

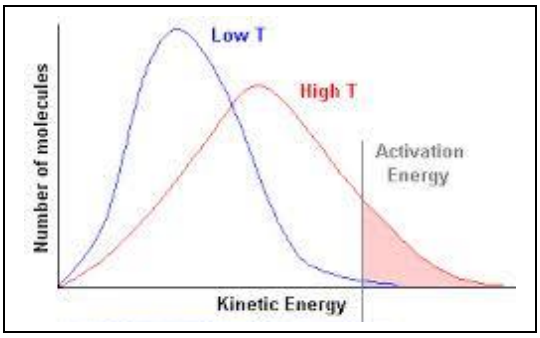
iii. The two reactions below are conducted under the same conditions.

$A_2(g) + B_2(g) \leftrightarrows 2AB(g)$ and

$2AB(g) \leftrightarrows A_2(g) + B_2(g)$

What can you say about the rates of the two reactions below?

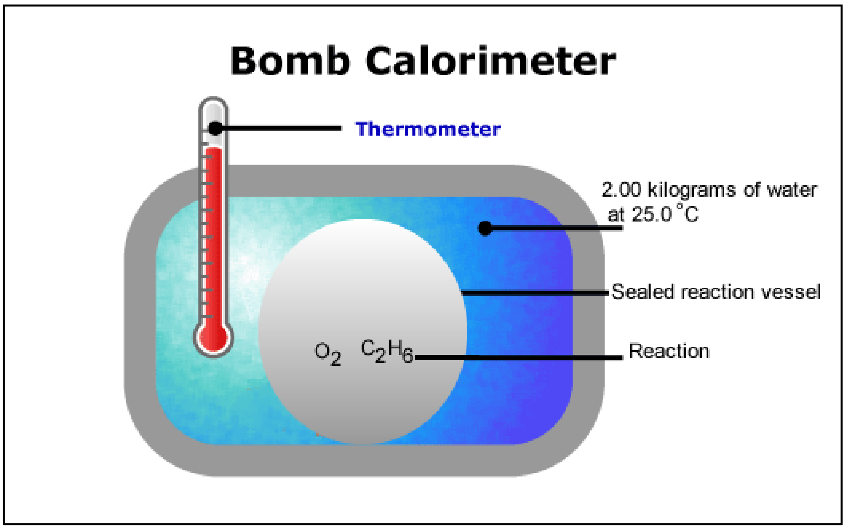
The forward reaction $A_2(g) + B_2(g) \leftrightarrows 2AB(g)$ has an activation energy of 80 $kJ.mol^{-}$ while the reverse reaction has an activation energy of 180 $kJ.mol^{-}$. The greater the activation energy the less number of molecules that will have this amount of energy and hence a slower rate of reaction results., as shown below



2) Ethane burns in oxygen according to the equation below

$$2C_2H_6(g) + 7O_2(g) \rightarrow 6H_2O(g) + 4CO_2(g)$$

A sample of 32.10 grams of oxygen is placed in a sealed container, as shown below, with excess ethane and ignited. The temperature of the water reached a maximum of 78.5 °C.



a) Calculate the mol of oxygen that reacted.

32.10 /32.0 = 1.00

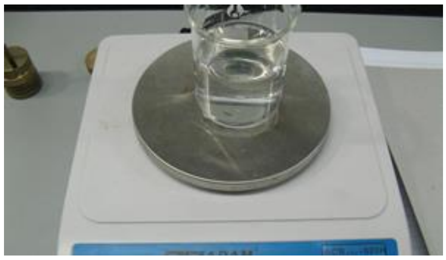
b) Calculate the amount of energy released, assuming no energy is lost from the system, using the information given in the question above.

E = 4.18 J/g/°C X mass X ΔT
⇨ = 4.18 X 2000 X 53.5 °C
⇨ = 447 kJ

c) Calculate the ΔH for the equation $2C_2H_6(g) + 7O_2(g) \rightarrow 6H_2O(g) + 4CO_2(g)$

Step 1 calculate the amount of energy released per mol of oxygen
447kJ / 1.00 = 447 kJ/ mol
Step 2 According to the equation 7 mol of oxygen gas is required.
ΔH = 7 X 447kJ = 3.13×10^3 kJ.

3) Hydrogen peroxide decomposes to form oxygen gas and water, according to the equation below.

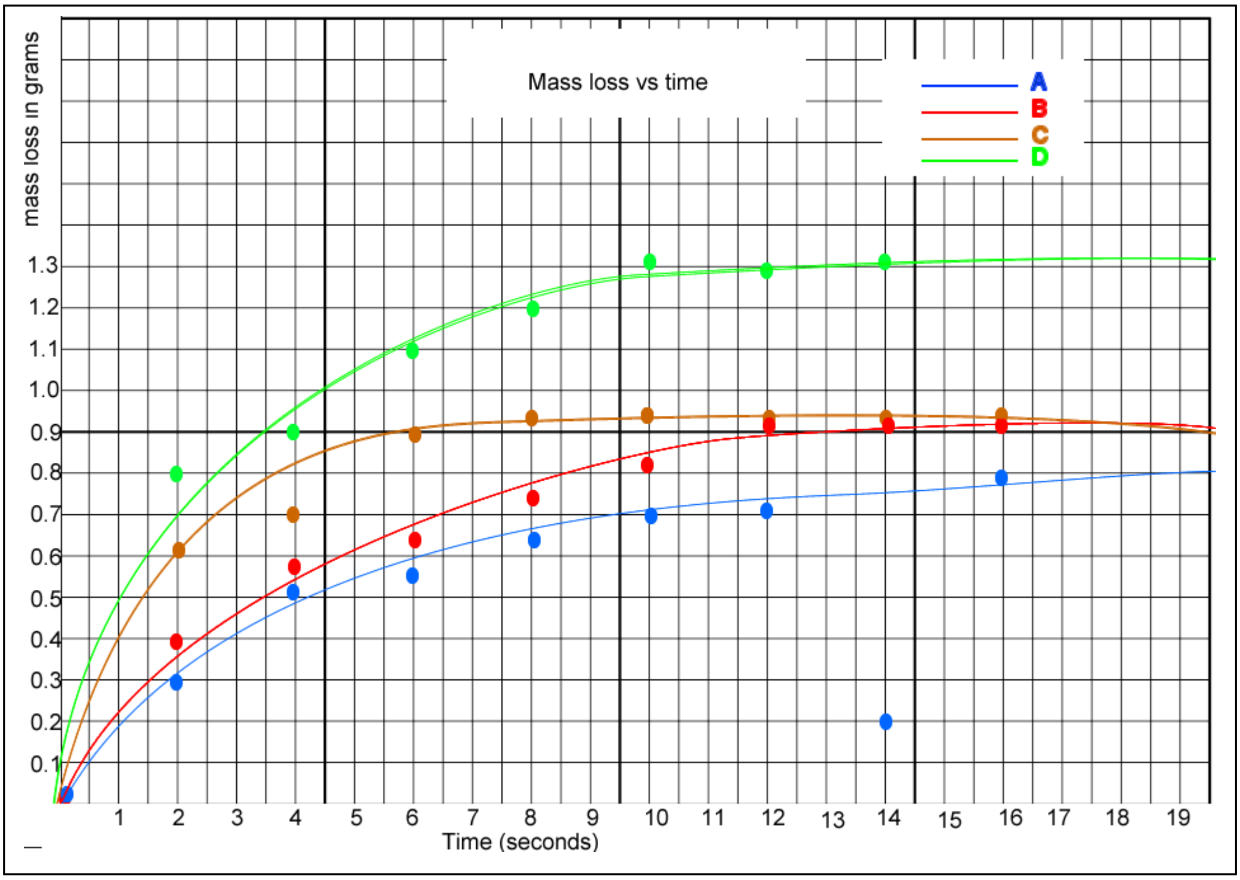


$$2H_2O_2(aq) \Rightarrow 2\,H_2O(l) + O_2(g)$$

In an investigation to measure the rate of decomposition of H_2O_2 was conducted by a student who set up the following experiment. He set up four beakers containing 100 mL of 30% v/v of H_2O_2 each. Beaker A was kept at 10 °C, while beaker B was kept at 20 °C, beaker C at 60 °C and beaker D at 90 °C. To start the reaction, a small amount of potassium iodide was placed in each beaker, the amount varied from beaker to beaker. The student recorded the following data for mass loss in grams (measured from the total starting mass) and the time, in seconds, from the start of the reaction. **<u>All data</u>** were accurately recorded in the table.

Time(sec)	***Reactants total mass loss (grams)***			
	Beaker B	Beaker A	Beaker C	Beaker **D**
0.00	0.00	0.00	0.00	0.00
2	0.40	0.30	0.60	0.80
4	0.58	0.50	0.70	0.90
6	0.64	0.56	0.90	1.10
8	0.75	0.66	0.93	1.20
10	0.82	0.70	0.93	1.31
12	0.92	0.74	0.93	1.30
14	0.92	0.20	0.92	1.29
16	0.92	0.80	0.93	1.30

a) Plot the graphs of the above data on one set of axes on the graphing grid below.



b) Formulate an appropriate hypothesis for this experiment.
An increase in temperature will increase the rate of a reaction.

c) Explain the results in terms of the particle -collision theory.
It is clear that the higher the temperature the steeper the graph of mass loss and hence the greater the rate of the reaction. At greater temperatures the average kinetic energy of particles increases and hence a greater number of molecules have the necessary activation energy to react.

d) What does the gradient of each graph indicate?

The rate of the forward reaction.

e) How would you improve this experiment?
The amount of potassium iodide added would have to be identical in all four beakers.

f) Give an explanation as to how beaker B may be different to the others.
A likely scenario is that more than 100 mL of H_2O_2 was added.

g) In which beaker is the reaction still proceeding at a slow rate after 18 seconds? Explain why.
Beaker A. Most likely the low temperature has slowed the reaction rate significantly.