151.90 g/mol

$$10\,FeSO_4 + 2\,KMnO_4 + 8\,H_2SO_4 \rightarrow 5\,Fe_2(SO_4)_3 + 2\,MnSO_4 + K_2SO_4 + 8\,H_2O$$

How many mL of 0.250M potassium permangenate are needed to react with 3.36 g of iron(II) sulfate?

---

1) Convert 3.36 grams iron(II) sulfate to moles. Use FORMULA WEIGHT.
2) Convert moles iron(II) sulfate to moles potassium permangenate. Use CHEMICAL EQUATION.
3) Convert moles potassium permangenate to solution volume. Use MOLARITY (0.250 M).

---

① $151.90\,g\ FeSO_4 = mol\ FeSO_4$ ② $10\ mol\ FeSO_4 = 2\ mol\ KMnO_4$

③ $0.250\ mol\ KMnO_4 = L$

$$3.36\,g\ FeSO_4 \times \underset{①}{\frac{mol\ FeSO_4}{151.90\,g\ FeSO_4}} \times \underset{②}{\frac{2\ mol\ KMnO_4}{10\ mol\ FeSO_4}} \times \underset{③}{\frac{L}{0.250\ mol\ KMnO_4}} = 0.0176958525\,L$$

Convert final answer to mL, since the problem specified that unit.

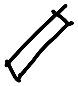
$mL = 10^{-3}\,L$

$$0.0176958525\,L \times \frac{mL}{10^{-3}\,L} = \boxed{17.7\ mL}\ \text{of}\ 0.250\ M\ KMnO_4$$

CONCEPT OF LIMITING REACTANT

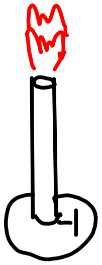
- When does a chemical reaction STOP?

$$2Mg(s) + O_2(g) \xrightarrow{\Delta} 2MgO(s)$$



Magnesium strip

+



Flame, oxygen from air

Magnesium oxide powder

- When does this reaction stop? When burned in open air, this reaction stops when all the MAGNESIUM STRIP is gone. We say that the magnesium is LIMITING.

- This reaction is controlled by the amount of available magnesium

- At the end of a chemical reaction, the LIMITING REACTANT will be completely consumed but there may be amount of OTHER reactants remaining. We do chemical calculations in part to minimize these "leftovers".

These are often called "excess" reactants, or reactants present "in excess"

LIMITING REACTANT CALCULATIONS

- To find the limiting reactant, calculate how much product would be produced from ALL given reactants. Whichever produces the SMALLEST amouht of product is the limiting reactant, and the smallest anount of product is the actual amount of product produced.

Example:

$$\overset{56.08}{\underline{CaO}(s)} + 3\,\overset{12.01}{\underline{C}}(s) \xrightarrow{\Delta} \overset{64.10}{CaC_2}(s) + CO(g)$$

<- Formula weights

If you start with 100. g of each reactant, how much calcium carbide would be produced?

---

① 56.08 g CaO = mol CaO ② mol CaO = mol CaC_2 ③ 64.10 g CaC_2 = mol CaC_2

$$100.\text{ g CaO} \times \frac{\text{mol CaO}}{56.08\text{ g CaO}} \times \frac{\text{mol CaC}_2}{\text{mol CaO}} \times \frac{64.10\text{ g CaC}_2}{\text{mol CaC}_2} = \boxed{114\text{ g CaC}_2}$$

---

① 12.01 g C = mol C ② 3 mol C = mol CaC_2 ③ 64.10 g CaC_2 = mol CaC_2

$$100.\text{ g C} \times \frac{\text{mol C}}{12.01\text{ g C}} \times \frac{\text{mol CaC}_2}{3\text{ mol C}} \times \frac{64.10\text{ g CaC}_2}{\text{mol CaC}_2} = 178\text{ g CaC}_2$$

The reaction stops when 114 grams of calcium carbide have been made. At that point, there is no more CaO left. We say that CaO is LIMITING and C is present IN EXCESS.

PERCENT YIELD

- Chemical reactions do not always go to completion! Things may happen that prevent the conversion of reactants to the desired/expected product!

① SIDE REACTIONS:

$$C + O_2 \longrightarrow CO_2$$

This reaction occurs when there is a large amount of oxygen available

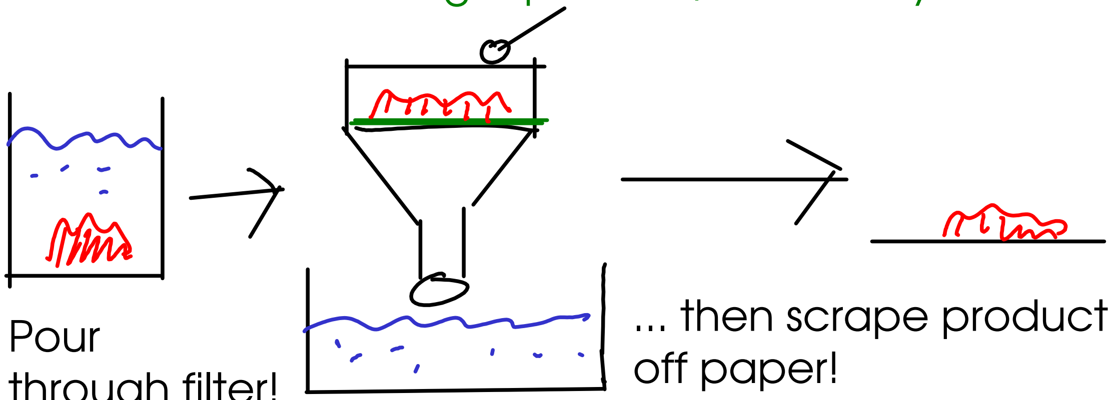
$$2C + O_2 \longrightarrow 2CO$$

... while this reaction is more favorable in low-oxygen environments!

... so in a low-oxygen environment, you may produce less carbon dioxide than expected!

② TRANSFER AND OTHER LOSSES

- When isolating a product, losses may occur in the process. Example: filtering



Pour through filter!

... then scrape product off paper!

During each step of this process, some amount of product loss will occur!

③ EQUILIBRIUM

- Reactions may reach an equilbrium between prodcuts and reactants. We'll talk more about this in CHM 111. The net results is that the reaction will appear to stop before all reactants have been consumed!

- All of these factors cause a chemical reaction to produce LESS product than calculated. For many reactions, this difference isn't significant. But for others, we need to report the PERCENT YIELD.

Determined EXPERIMENTALLY!

$$\text{PERCENT YIELD} = \frac{\text{ACTUAL YIELD}}{\text{THEORETICAL YIELD}} \times 100\,\%$$

Calculated based on the limiting reactant. (The chemical calculations you've done up to now have been theoretical yields!)

... the percent yield of a reaction can never be greater than 100% due to conservation of mass! If you determine that a percent yield is greater than 100%, then you've made a mistake somewhere - either in a calculation or in the experiment itself!

78.114 g/mol 123.111 g/mol <- Formula weights

22.4 g 31.6 g ACTUAL

$$C_6H_6 + HNO_3 \longrightarrow C_6H_5NO_2 + H_2O$$

benzene nitric acid nitrobenzene

22.4 grams of benzene are reacted with excess nitric acid. If 31.6 grams of nitrobenzene are collected from the reaction, what is the percent yield?

Calculate the THEORETICAL YIELD of nitrobenzene that could be made from 22.4 grams of benzene (the starting material!)

① $78.114\,g\ C_6H_6 = mol\ C_6H_6$ ② $mol\ C_6H_6 = mol\ C_6H_5NO_2$

③ $123.111\,g\ C_6H_5NO_2 = mol\ C_6H_5NO_2$

$$22.4\,g\ C_6H_6 \times \frac{mol\ C_6H_6}{78.114\,g\ C_6H_6} \times \frac{mol\ C_6H_5NO_2}{mol\ C_6H_6} \times \frac{123.111\,g\ C_6H_5NO_2}{mol\ C_6H_5NO_2} = 35.3\,g\ C_6H_5NO_2$$

(theoretical)

$$\%\,yield = \frac{actual}{theor.} \times 100 = \frac{31.6g}{35.3g} \times 100 = \boxed{89.5\,\%}$$

ENERGY

- thermodynamics: the study of energy transfer

Conservation of energy: Energy may change form, but the overall amount of energy remains constant. "first law of thermodynamics"

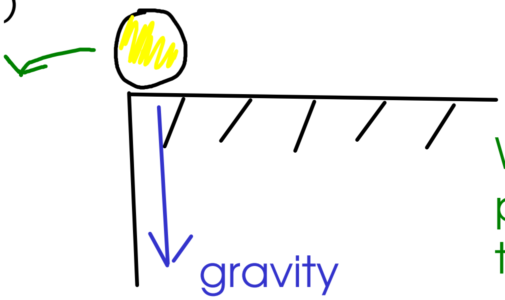
- ... but what IS energy?
  - energy is the ability to do "work"

motion of matter

Kinds of energy?

- Kinetic energy: energy of matter in motion $E_K = \frac{1}{2} \overset{\text{mass}}{m} \underset{\text{velocity}}{v}^2$

- Potential energy: energy of matter that is being acted on by a field of force (like gravity)



When the ball falls, its potential energy is converted to kinetic!

- What sort of energy concerns chemists? Energy that is absorbed or released during chemical reactions.

- Energy can be stored in chemicals ... molecules and atoms.

INTERNAL ENERGY: "U"

related to the kinetic and potential energy of atoms, molecules, and their component parts.

- We measure energy transfer ... which is called HEAT. (HEAT is the flow of energy from an area of higher temperature to an area of lower temperature)

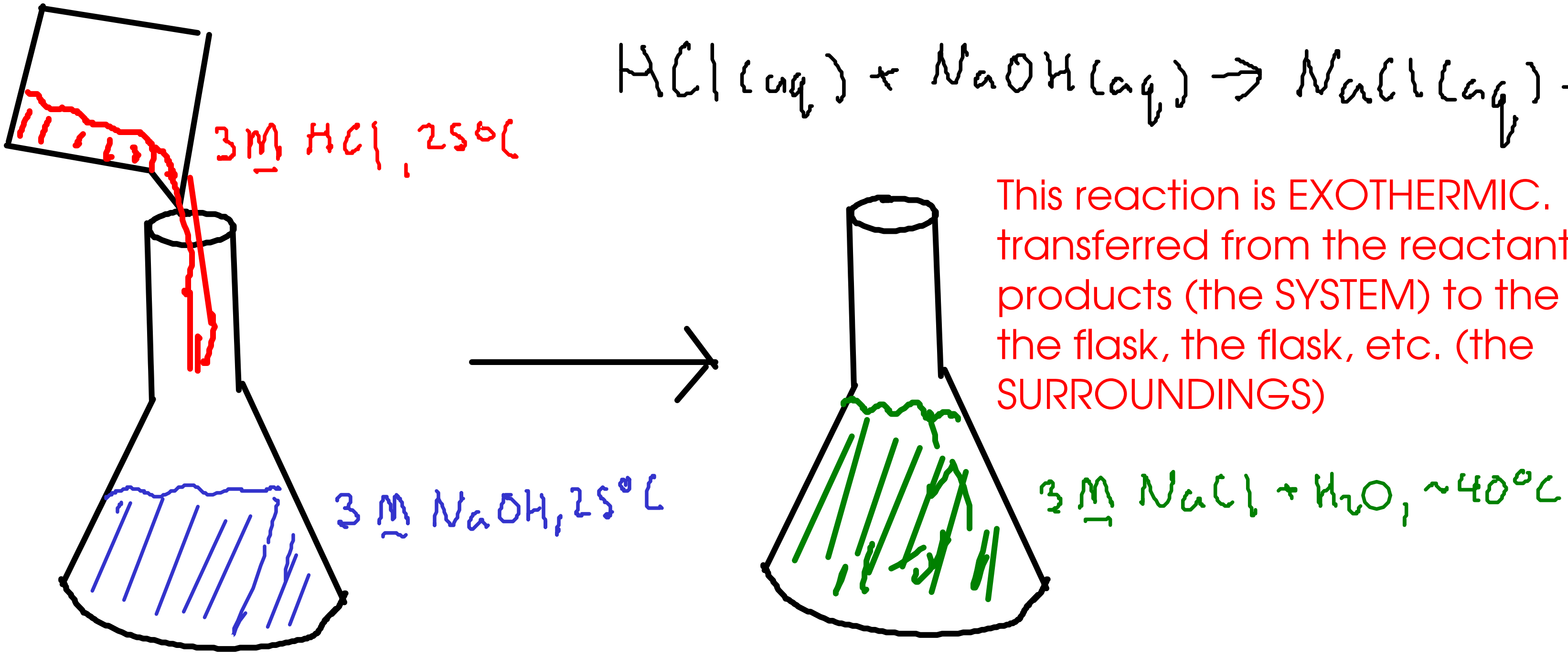
Q : heat

SYSTEM: the object or material under study

SURROUNDINGS: everything else

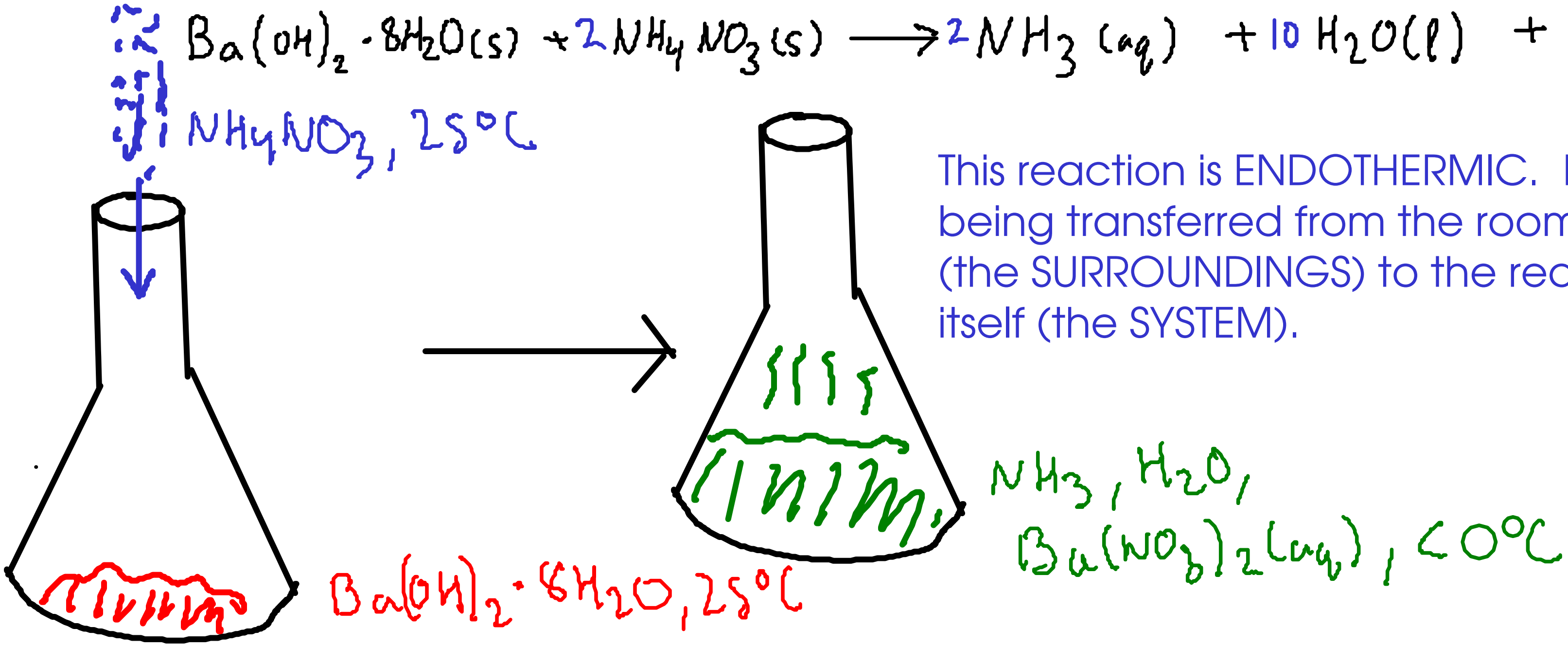
Type of process	Energy is ...	Sign of Q	Temp of SURROUNDINGS ...
ENDOTHERMIC	transferred from SURROUNDINGS to SYSTEM	+	decreases
EXOTHERMIC	transferred from SYSTEM to SURROUNDINGS	—	increases

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$



This reaction is EXOTHERMIC. Energy is transferred from the reactants and products (the SYSTEM) to the water in the flask, the flask, etc. (the SURROUNDINGS)

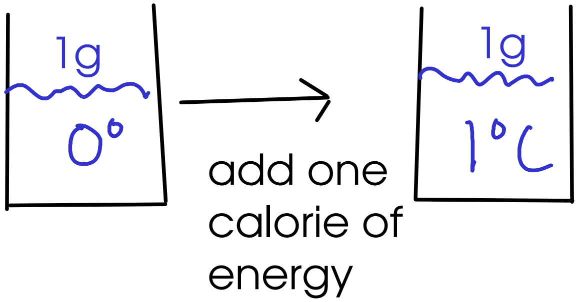
$Ba(OH)_2 \cdot 8H_2O(s) + 2NH_4NO_3(s) \rightarrow 2NH_3(aq) + 10H_2O(l) + Ba(NO_3)_2(aq)$



This reaction is ENDOTHERMIC. Energy is being transferred from the room/flask/etc. (the SURROUNDINGS) to the reaction itself (the SYSTEM).

ENERGY UNITS

- calorie (cal): the amount of energy required to change the temperature of one gram of water by one degree Celsius (or Kelvin)



1g ≈ 1mL for water

- Calories in food? The "Calorie" that is given on American food labels is actually the kilocalorie (kcal)

- Joule (J): SI unit for energy. It's defined based on the equation for kinetic energy.

$$1\,J = 1\,\frac{kg\,m^2}{s^2}$$, from

$$E_K = \frac{1}{2}mv^2$$

kinetic energy

mass

velocity

$$\boxed{4.184\ J = 1\ cal}$$

- the Joule is a small unit. For most reactions at lab scale, we'll use kilojoules (kJ).

SPECIFIC HEAT AND HEAT CAPACITY

- a measured quantity. The amount of energy required to change the temperature of one gram of a particular substance by one degree Celsius.

- Specific heat information for common substances is readily available. For water,

$$4.184\ \frac{J}{g\,{}^{\circ}C} \quad \underset{=}{\text{or}} \quad 1.000\ \frac{cal}{g\,{}^{\circ}C}$$

$$Q = m \times S \times \Delta T$$

m = mass
s = specific heat
ΔT = Tfinal - Tinitial

This is ALWAYS final temp minus initial temp!

- For objects, like reaction vessels, you might know the HEAT CAPACITY, which is the amount of energy required to change the temperature of an object by one degree Celsius

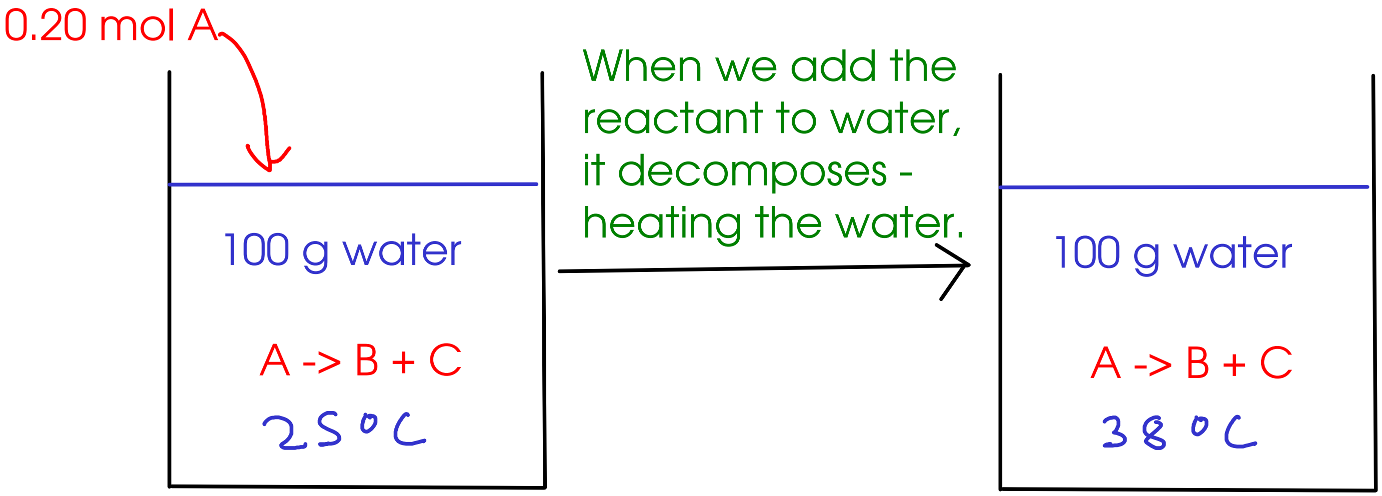
Units: $J/{}^{\circ}C$ or $cal/{}^{\circ}C$

$$Q = C \times \Delta T$$

c = heat capacity

CALORIMETRY

- the measurement of heat. But how do we measure heat?



When we add the reactant to water, it decomposes - heating the water.

... what is Q for this reaction?

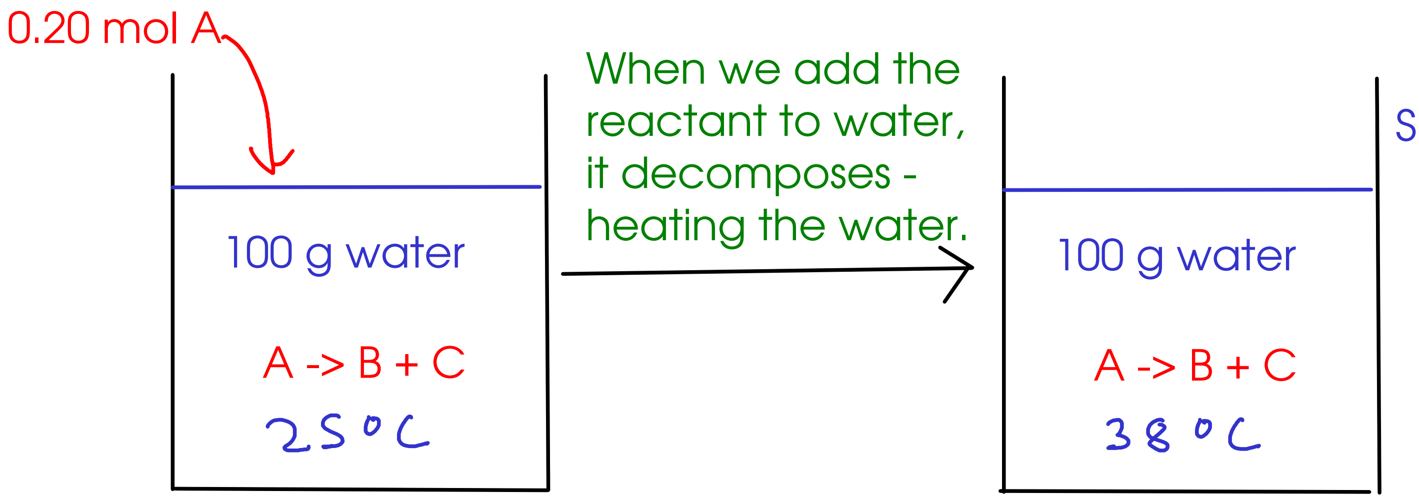
Assuming that no heat is lost from the water to the surrounding air,

$$\underbrace{Q_r}_{\text{reaction}} + \underbrace{Q_w}_{\text{water}} = 0$$

Conservation of energy. The terms add to zero because they have opposite signs.

... if we knew something about the WATER, we could use that to find the heat of the REACTION!

We can look up the water's SPECIFIC HEAT and use it to relate the temperature change of the water to Q.



When we add the reactant to water, it decomposes - heating the water.

Specific heat of water:

$4.184 \; J/g°C$

$$Q_r + Q_w = 0$$

$$Q_w = m_w \times S_w \times \Delta T_w = (100 g)(4,184 \; J/g°C)(38°C - 25°C)$$
$$= 5439,2 J$$

$$Q_r + Q_w = 0, \quad Q_r + 5439.2 J = 0, \quad Q_r = -5439.2 J$$

To report the energy change in this reaction to others, we should express it in terms of heat transfer per mole of something. A different amount of reactant would have a different Q

$$Q_{rxn} = \frac{Q_r}{\text{moles A}} = \frac{-5439.2 J}{0.20 \text{ mol A}} = -27000 \frac{J}{\text{mol A}} = \boxed{-27 \frac{kJ}{\text{mol A}}}$$

This kind of number is usally called a "heat of reaction" ...

One problem ...

PATH. The amount of energy required for a process depends on how the process is carried out.

Example: Driving from Florence to Columbia. How much energy is required? (gas)

Jeep Cherokee vs Toyota Prius. The Jeep will use much more fuel than the Prius even though they start and end from exactly the same place. So the fuel usage is what we call a PATH FUNCTION, while the location is a STATE FUNCTION.

- so the heat of reaction depends on how the reaction is done.
  - we need (for reporting) some kind of standard condition. At constant pressure, we can define a state function called ENTHALPY (H)

$$H = U + PV$$

$$\Delta H = Q_{\text{constant pressure}}$$

... we record the "enthalpy change of reaction" in our data books.

$$\Delta H_r$$

SINCE the enthalpy change does NOT depend on path, this means that we can use standard values for enthalpy to predict the heat change in reactions that we have not tested in a calorimeter.