N38 – Energy of Reactions

Target: I can perform various calculations involving the energy changes during chemical reactions.

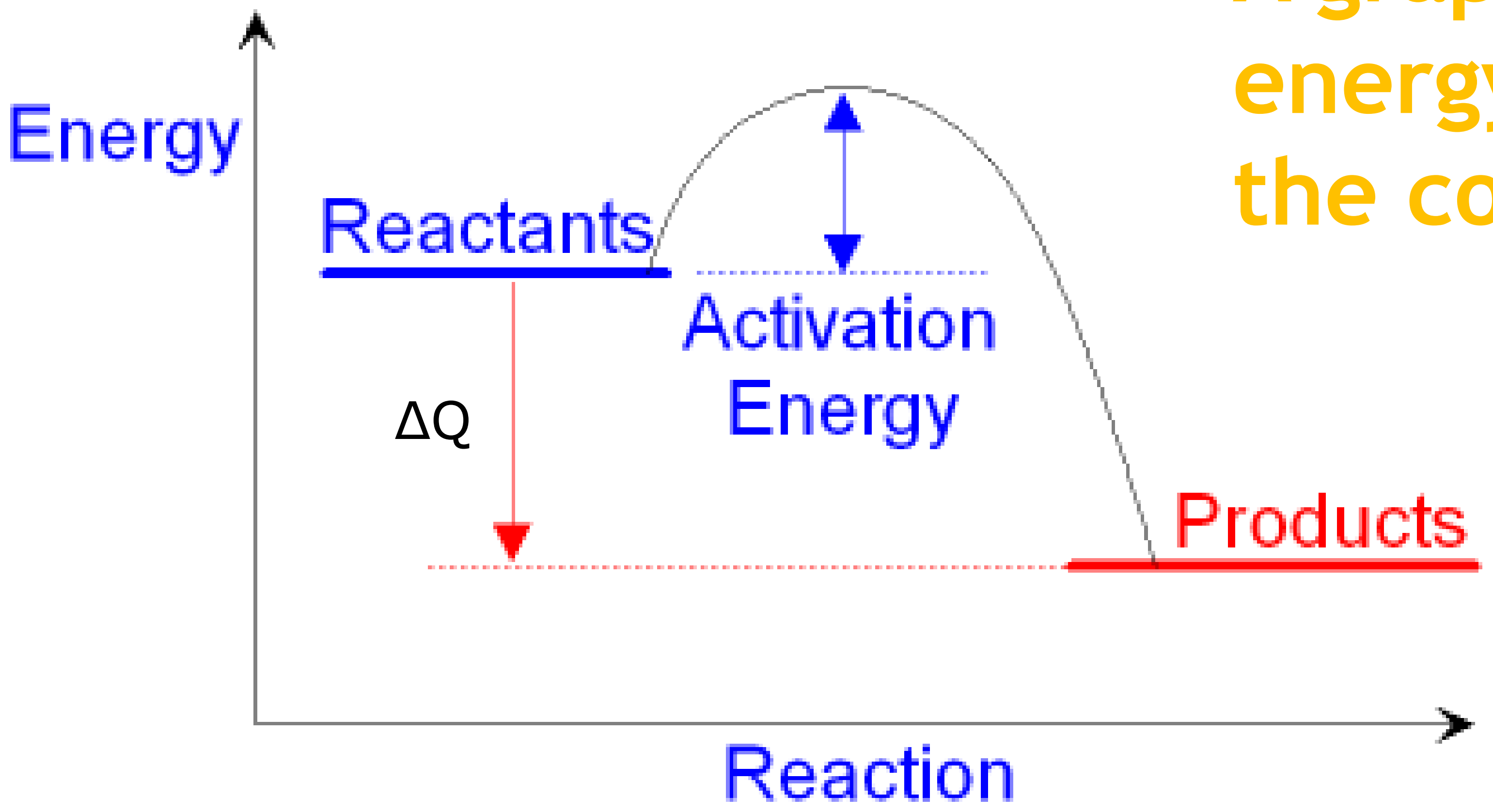
Link to YouTube Presentation: https://youtu.be/A7z5ixKMBQs

N38 – Energy of Reactions

1. Reaction Diagrams
2. Molar Heat Capacities
3. Heat of Reaction
4. Heat of Formation
5. Bond Energy
6. Hess's Law

1. Reaction Diagrams

Reaction Diagrams

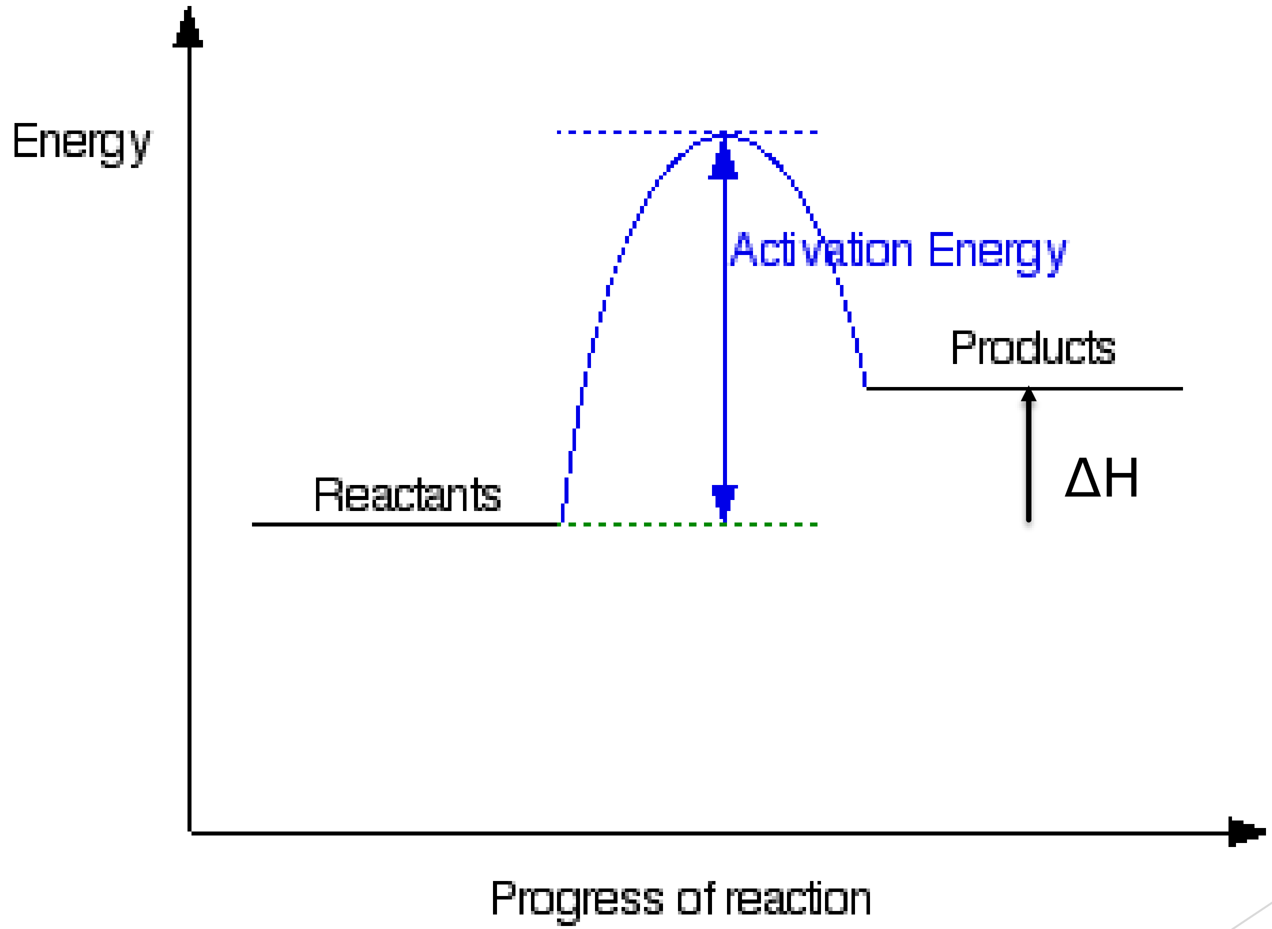


A graph representing energy changes during the course of a reaction

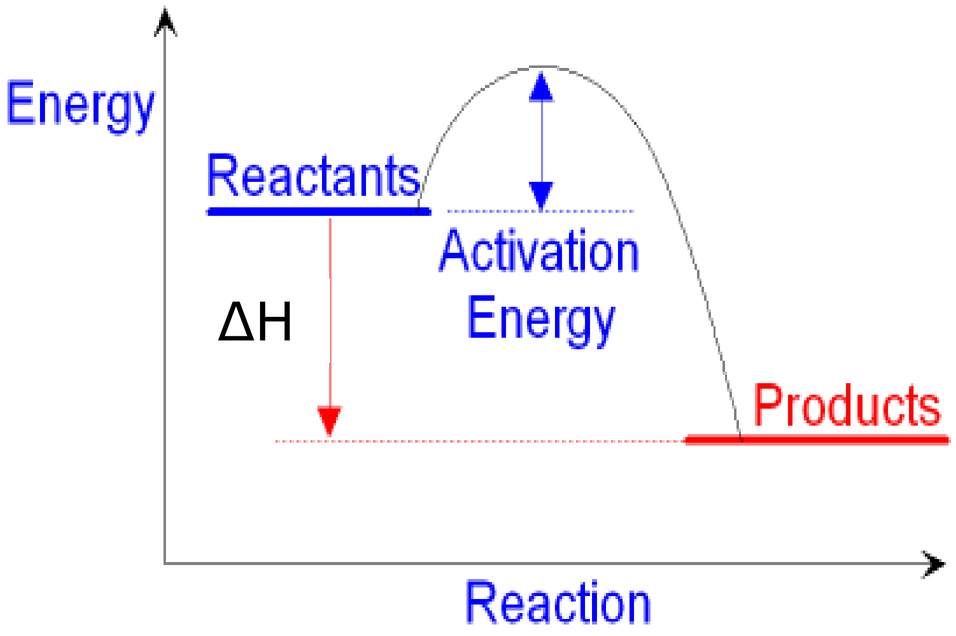
*Is this reaction endo or exothermic? How do you know???

Reaction Diagrams

Draw and label an ENDOTHERMIC reaction diagram.



But what is "Activation Energy?"



<u>Activation energy:</u>
the smallest amount of energy required for molecules to be "activated" in order to undergo a specific chemical change

- ***Speed them up to hit hard enough***
- ***Proper orientation to collide in the right spot***

2.
Molar
Heat Capacity

Molar Heat Capacity

Energy required to raise the temperature of one <u>MOLE</u> of a substance one degree. Similar to Specific Heat Capacity but uses moles instead of grams!

$$Q = nC\Delta T$$

**If you make sure your units cancel, this is easy!!*

Molar Heat Capacity

A sample of barium chloride is increased in temperature by 3.8°C when the sample absorbed 2.4×10^2 J of heat energy. Calculate the number of moles of barium chloride if its molar heat capacity is 75.1 J/K•mol.

$Q = nC\Delta T$

$$2.4 \times 10^2 \text{ J} = n\left(75.1 \frac{J}{K \bullet mol}\right)(3.8 \text{ K})$$

$$n = 0.84 \text{ mol}$$

Do we care that it is in moles and Kelvins? No! Does that change the concept of "plug and chug" and "cancel units" ? No!

3. Heat of Reaction

Heat of Reactions

Amount of energy change involved in a reaction.
Sometimes exo, sometimes endo.

Lots of different ways to calculate or measure it.

$2Al + Fe_2O_3 \rightarrow 2Fe + Al_2O_3 \qquad \Delta H_{rxn} = -851.5 kJ/mol_{rxn}$

(Remember, ΔH is basically Q)

ΔH negative → energy released → exothermic → **product!**

ΔH positive → energy absorbed → endothermic → **reactant!**

Heat of Reactions

Amount of energy change involved in a reaction.
Sometimes exo, sometimes endo.

Lots of different ways to calculate or measure it.

$$2Al + Fe_2O_3 \rightarrow 2Fe + Al_2O_3 + \text{energy}$$

$$2Al + Fe_2O_3 \rightarrow 2Fe + Al_2O_3 + 851.5\text{kJ}$$

Energy released! Product! Exothermic!

Heat of Reactions

Amount of energy change involved in a reaction.
Sometimes exo, sometimes endo.

Lots of different ways to calculate or measure it.

$$2Fe + Al_2O_3 + \text{energy} \rightarrow 2Al + Fe_2O_3$$

$$2Fe + Al_2O_3 + 851.5\text{kJ} \rightarrow 2Al + Fe_2O_3$$

Energy absorbed! Reactant! Endothermic!

Heat of Reactions per mole

Sometimes you want it per mole of a certain substance.

Just take mole ratios into account!

$$2Al + Fe_2O_3 \rightarrow 2Fe + Al_2O_3 \qquad \Delta H_{rxn} = -851.5 kJ/mol_{rxn}$$

$$\frac{-851.5kJ}{1\ rxn} \cdot \frac{1\ rxn}{2\ mol\ Al} = -425.75\ \frac{kJ}{mol\ Al}$$

Example Question

Calculate the energy released when 135g of aluminum is reacted in the below equation.

$$2Al + Fe_2O_3 \rightarrow 2Fe + Al_2O_3 \qquad \Delta H_{rxn} = -851.5 kJ/mol_{rxn}$$

$$\frac{135\ g\ Al}{}$$

Example Question

Calculate the energy released when 135g of aluminum is reacted in the below equation.

$$2Al + Fe_2O_3 \rightarrow 2Fe + Al_2O_3 \qquad \Delta H_{rxn} = -851.5 kJ/mol_{rxn}$$

135 g Al	1 mol Al	1 rxn	-851.5 kJ	= -2130.3 kJ
	26.98g Al	2 mol Al	1 rxn	

4.
Heat of Formation

Heat of Formation

If you know how much energy it takes to form each substance in a reaction, you can calculate the Heat of Reaction!

$$\Delta \boldsymbol{H}^{\circ} = \boldsymbol{\Sigma n \Delta H^{\circ}_{f}(products)} - \boldsymbol{\Sigma n \Delta H^{\circ}_{f}}(\text{reactants})$$

- Σ means sum.
- n is the coefficient of the reaction.
- values will be given to you in a chart.

Calculating Heat of Rxn from Heats of Formation

Calculate ΔH for the combustion of methane, CH_4

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$$

$$\Delta H^\circ = \Sigma n\Delta H_f^\circ(products) - \Sigma n\Delta H_f^\circ(\text{reactants})$$

Substance	ΔH_f (kJ)
CH_4	-74.80
O_2	0
CO_2	-393.50
H_2O	-285.83

ΔH_{rxn} = [-393.50kJ + 2(-285.83kJ)] – [-74.80kJ +2(0kJ)]

ΔH_{rxn} = -890.36 kJ/mol_{rxn}

Ethanol is used as an additive in many fuels today.
What is ΔH^{o}_{rxn} (kJ) for the combustion of ethanol?

$2\ C_2H_5OH_{(l)} + 6\ O_{2\,(g)} \rightarrow 4\ CO_{2\,(g)} + 6\ H_2O_{(l)}$

Formula	ΔH^{o}_{f}
C_2H_5OH (*l*)	–277.6
CO_2 (*g*)	–393.5
H_2O (*g*)	–241.8
H_2O (*l*)	–285.8

- A - 401.7
- B + 401.7
- C - 2469
- D + 2734
- E - 2734

Ethanol is used as an additive in many fuels today.

What is ΔH^o_{rxn} (kJ) for the combustion of ethanol?

$2\ C_2H_5OH_{(l)} + 6\ O_{2\ (g)} \rightarrow 4\ CO_{2\ (g)} + 6\ H_2O_{(l)}$

Formula	**ΔH^o_f**
$C_2H_5OH\ (l)$	–277.6
$CO_2\ (g)$	–393.5
$H_2O\ (g)$	–241.8
$H_2O\ (l)$	–285.8

A - 401.7

B + 401.7

C - 2469

D + 2734

E **- 2734**

$\Delta H^o_{rxn} =$ *Products – Reactants*

$$[4(-393.5) + 6(-285.8)] - [2(-277.6) + 6(0)]$$

= **-2734 kJ/mol$_{rxn}$**

5. Bond Energy

Bond Energy

It TAKES energy to break a bond – ENDO

- Otherwise they would just break by themselves!

Energy is RELEASED when a new bond forms – EXO

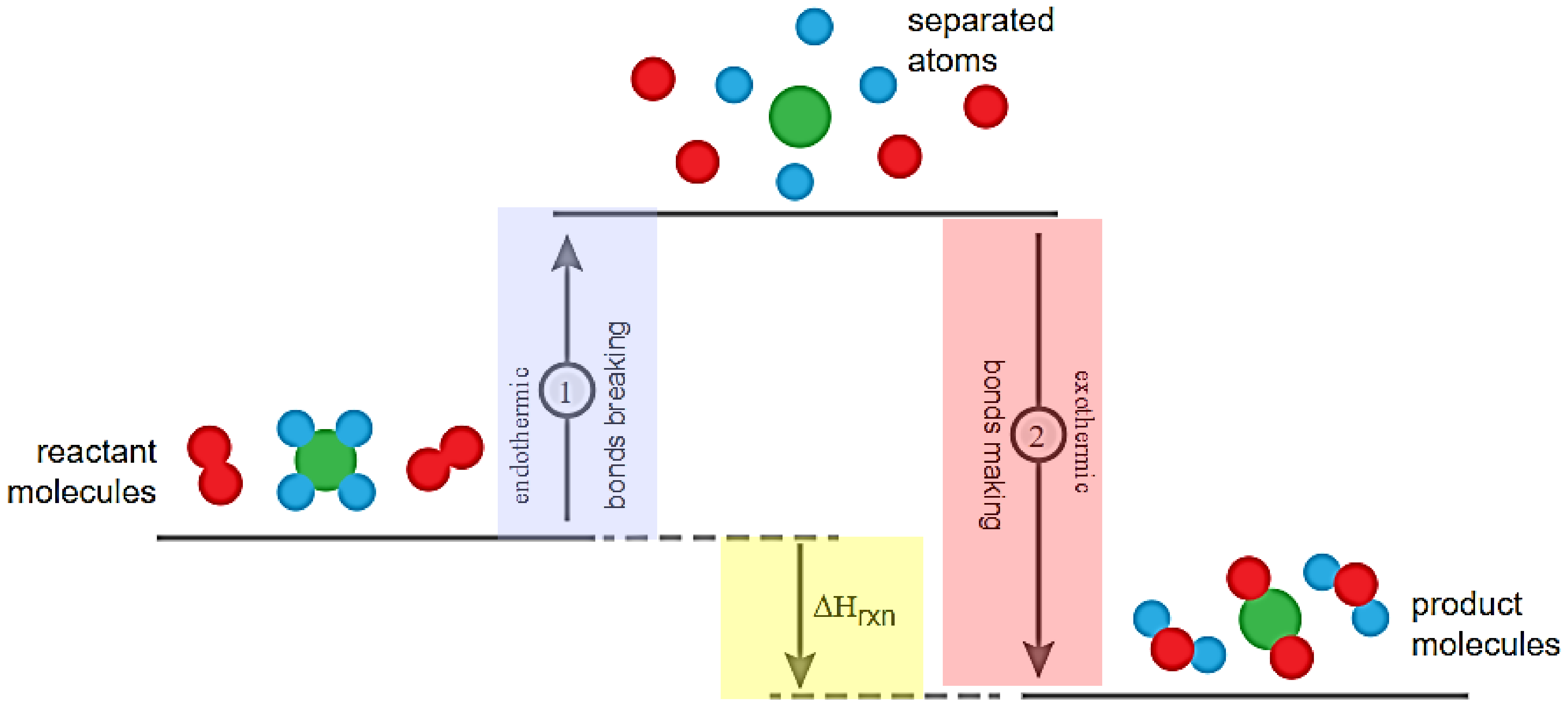
- If the new bond isn't more stable, lower energy, it wouldn't want to form!

Bond Energy

If you know how much energy it takes to break a bond, and how much energy is released when new bonds form, you can add it all up to figure out the Heat of Reaction based on the bonds that are changing during the reaction!

- Values will be given to you in a chart
- Values in the table are *averages*

Bond Energy


separated
atoms
reactant
molecules
endothermic
1
bonds breaking
bonds making
2
exothermic
ΔHrxn
product
molecules

Use Bond Energy Chart

Action	Algebraic Sign	How to Remember
Break a Bond	+	Takes to Break
Form a Bond	-	Free to Form

$2H_2 + O_2 \rightarrow 2H_2O$

You have to break: 2 H-H bond and 1 O=O bond
You have to form: 4 H-O bonds

2(436) + (498) + 4(- 463)= **- 482 kJ/mol** (exo)

Single Bond Energies (kJ/mol of bonds)

	H	C	N	O	S	F	Cl	Br	I
H	436								
C	413	346							
N	391	305	163						
O	463	358	201	146					
S	347	272	—	—	226				
F	565	485	283	190	284	155			
Cl	432	339	192	218	255	253	242		
Br	366	285	—	201	217	249	216	193	
I	299	213	—	201	—	278	208	175	151

Multiple Bond Energies (kJ/mol of bonds)

C=C 602	C=N 615	C=O 799
C≡C 835	C≡N 887	C≡O 1072
N=N 418	N=O 607	
N≡N 945	O=O 498	

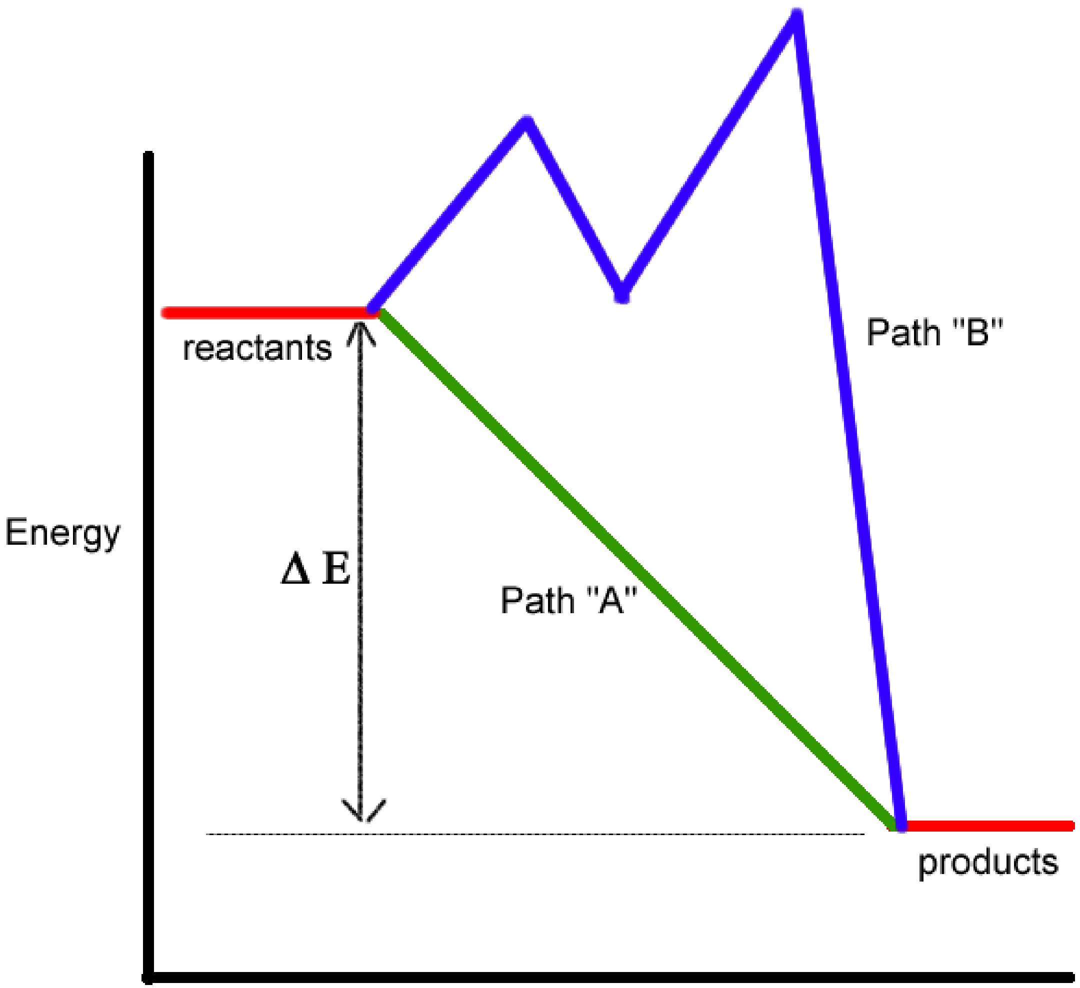
6. Hess's Law

Hess's Law

Path A – Mrs. Farmer cleaning the house.

Path B – Mr. Farmer cleaning the house.

Regardless of the path taken, you still get to the same place.

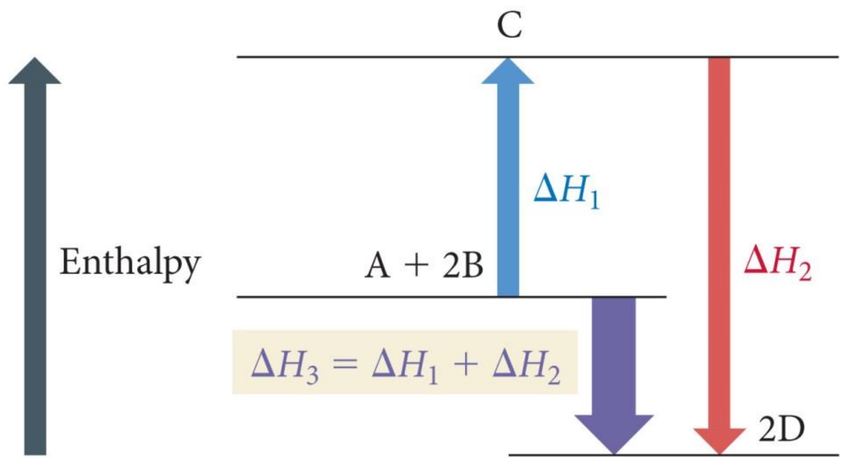


Although Path B drives Mrs. Farmer bonkers – Ha!

Hess's Law

"In going from a particular set of reactants to a particular set of products, the change in enthalpy is the same whether the reaction takes place in one step or a series of steps."

Hess's Law
The change in enthalpy for a stepwise process is the sum of the enthalpy changes of the steps.



Relationships Involving ΔH_{rxn}

Multiplying Rxn by a # to Change Coefficients

ΔH_{rxn} is multiplied by that factor.

- Because ΔH_{rxn} is *extensive* – depends on the amount of substance

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H = -393.5$ kJ

2 $C(s)$ + **2** $O_2(g) \rightarrow$ **2** $CO_2(g)$ $\Delta H =$ **2 x** $(-393.5$ kJ$) = -787.0$ kJ.

Reversing a rxn to flip which side the products/reactants are on

Flip the sign of ΔH, if positive now negative, if negative, now positive

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H = -393.5$ kJ

$CO_2(g) \rightarrow C(s) + O_2(g)$ $\Delta H =$ **−** $(-393.5) =$ **+** 393.5 kJ

Hess's Law Example Problem #1

Calculate ΔH for the combustion of methane, CH_4:

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

#	Reaction	ΔHº
1	$C + 2H_2 \rightarrow CH_4$	-74.80 kJ
2	$C + O_2 \rightarrow CO_2$	-393.50 kJ
3	$H_2 + ½ O_2 \rightarrow H_2O$	-285.83 kJ

Step #1:

CH_4 must appear on the reactant side, so we reverse reaction #1 and change the sign on ΔH.

- rxn 1 **$CH_4 \rightarrow C + 2H_2$** **-(-74.80 kJ)**

Hess's Law Example Problem #1

Calculate ΔH for the combustion of methane, CH_4:

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

#	Reaction	ΔHº
1	$C + 2H_2 \rightarrow CH_4$	-74.80 kJ
2	$C + O_2 \rightarrow CO_2$	-393.50 kJ
3	$H_2 + ½\ O_2 \rightarrow H_2O$	-285.83 kJ

Step #2:

Keep reaction #2 unchanged, because CO_2 belongs on the product side

- rxn 1 **$CH_4 \rightarrow C + 2H_2$** **-(-74.80 kJ)**

rxn 2 **$C + O_2 \rightarrow CO_2$** **-393.50 kJ**

Hess's Law Example Problem #1

Calculate ΔH for the combustion of methane, CH_4:

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

#	Reaction	ΔHº
1	$C + 2H_2 \rightarrow CH_4$	-74.80 kJ
2	$C + O_2 \rightarrow CO_2$	-393.50 kJ
3	$H_2 + ½ O_2 \rightarrow H_2O$	-285.83 kJ

Step #3:

Use reaction #3 to get water as a product, but multiply it by 2 since you have 2 H_2O

- rxn 1	**$CH_4 \rightarrow C + 2H_2$**	**-(-74.80 kJ)**
rxn 2	**$C + O_2 \rightarrow CO_2$**	**-393.50 kJ**
2 x rxn 3	**$2H_2 + O_2 \rightarrow 2H_2O$**	**2 x (-285.83 kJ)**

Hess's Law Example Problem #1

Calculate ΔH for the combustion of methane, CH_4:

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

#	Reaction	ΔHº
1	$C + 2H_2 \rightarrow CH_4$	-74.80 kJ
2	$C + O_2 \rightarrow CO_2$	-393.50 kJ
3	$H_2 + ½ O_2 \rightarrow H_2O$	-285.83 kJ

Step #4:

Cross out things that show up on both sides, then sum up your ΔH values

- rxn 1	$CH_4 \rightarrow C + 2H_2$	-(-74.80 kJ)
rxn 2	$C + O_2 \rightarrow CO_2$	-393.50 kJ
2 x rxn 3	$2H_2 + O_2 \rightarrow 2H_2O$	2 x (-285.83 kJ)
	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$	-890.36 kJ

Hess's Law Example Problem #2

$2\ NOCl\ (g) \rightarrow N_2\ (g) + O_2\ (g) + Cl_2\ (g) \quad \Delta H = \ ?$

Rxn #1) $\frac{1}{2}\ N_2\ (g) + \frac{1}{2}\ O_2\ (g) \rightarrow NO\ (g)$	$\Delta H =$	90.3 kJ
Rxn #2) $NO\ (g) + \frac{1}{2}\ Cl_2\ (g) \rightarrow NOCl\ (g)$	$\Delta H =$	–38.6 kJ

A) -51.7 kJ

B) 51.7 kJ

C) -103.4 kJ

D) 103.4 kJ

E) 142.0 kJ

Hess's Law Example Problem #2

$2\ NOCl\ (g) \rightarrow N_2\ (g) + O_2\ (g) + Cl_2\ (g)$ $\Delta H =$?

Rxn #1) $\frac{1}{2}\ N_2\ (g) + \frac{1}{2}\ O_2\ (g) \rightarrow NO\ (g)$	$\Delta H =$	90.3 kJ
Rxn #2) $NO\ (g) + \frac{1}{2}\ Cl_2\ (g) \rightarrow NOCl\ (g)$	$\Delta H =$	–38.6 kJ

A -51.7 kJ

B 51.7 kJ

C **-103.4 kJ**

D 103.4 kJ

E 142.0 kJ

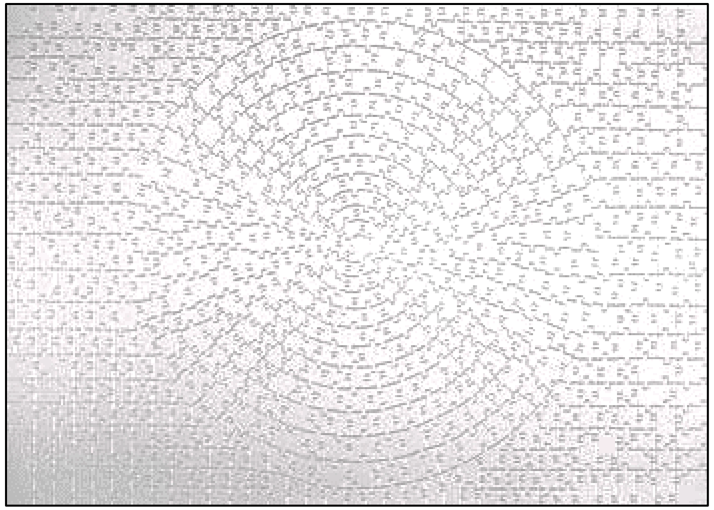
Rxn #	How to change it	Rxn	ΔH
2	- and x 2	$2\ NOCl \rightarrow$ ~~2NO~~ $+ Cl_2$	-2 (-38.6)
1	- and x 2	~~2NO~~ $\rightarrow N_2 + O_2$	- 2 (90.3)
		$2NOCl \rightarrow N_2 + O_2 + Cl_2$	**-103.4 kJ**

Its just a puzzle!

Sometimes it's a really hard puzzle… but it's still just a puzzle!

All the pieces are there, you just have to figure out how to put them together…

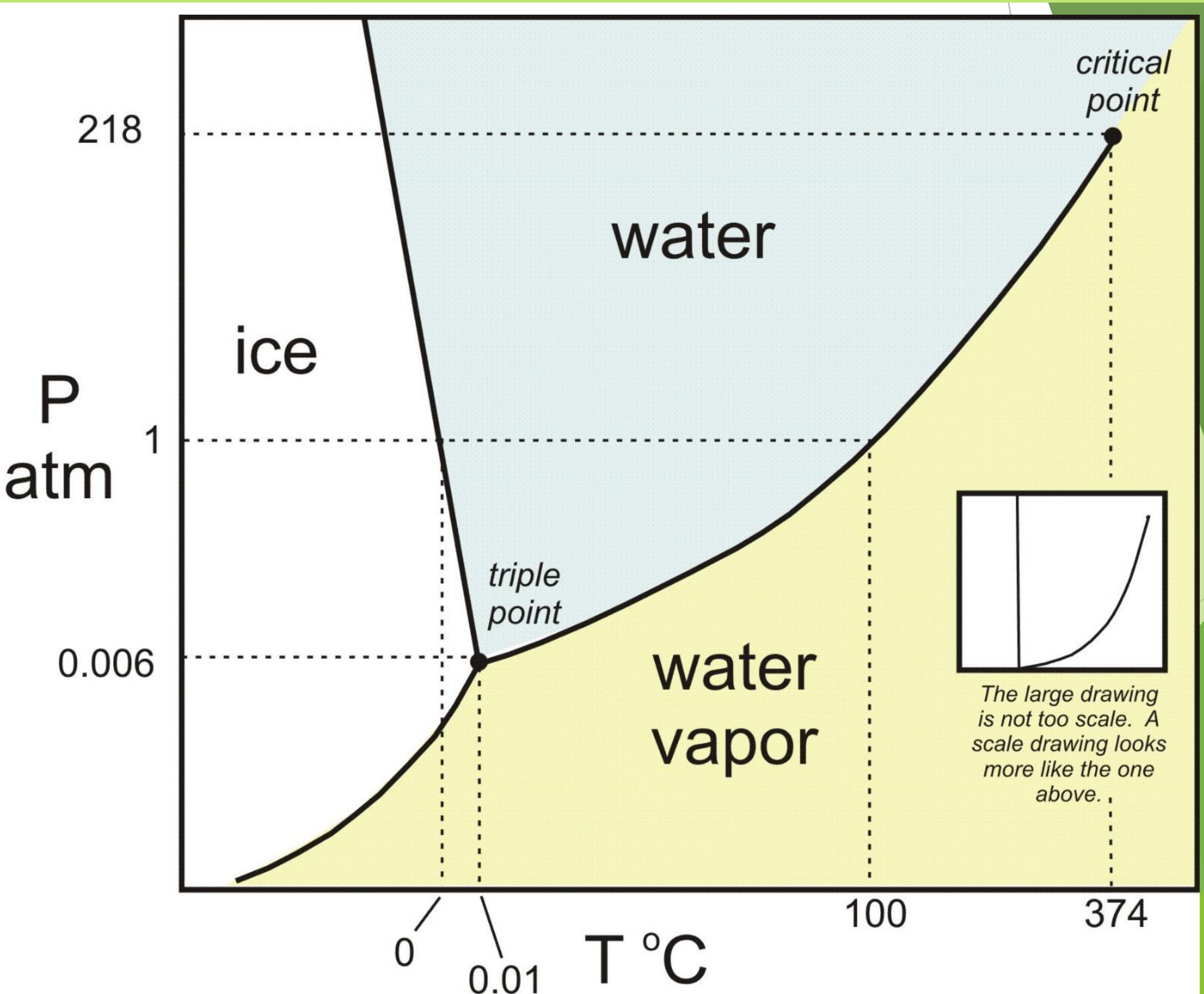
Unfortunately no real "tricks" for how to figure out which parts to put together.



Old stuff after this slide. You do not have to go over this stuff. BUT if you plan to take AP Chem you may want to take a peek! ☺

Phase Diagrams

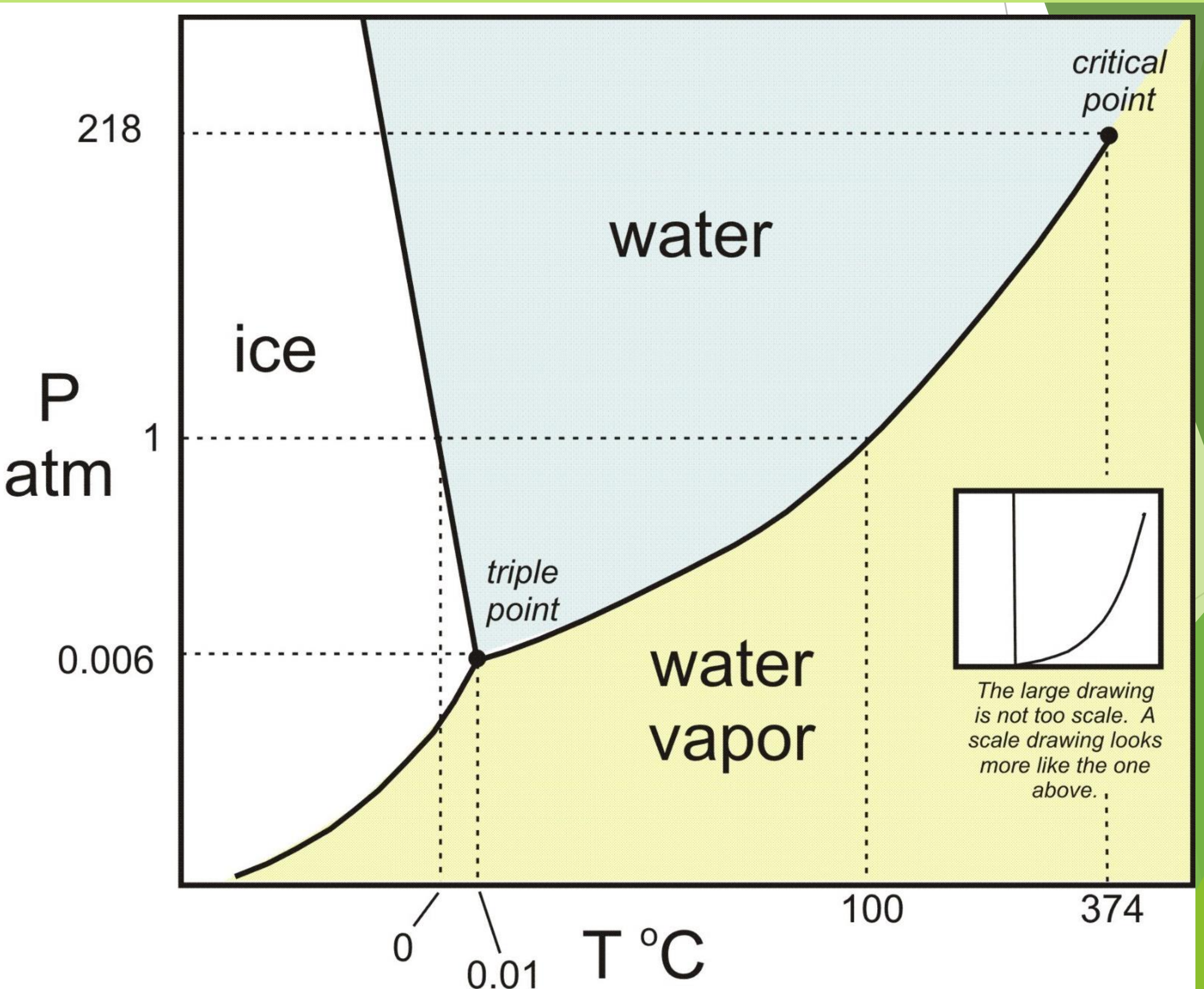
A graph representing the phases of a substance at a given temperature and pressure



Phase Diagrams

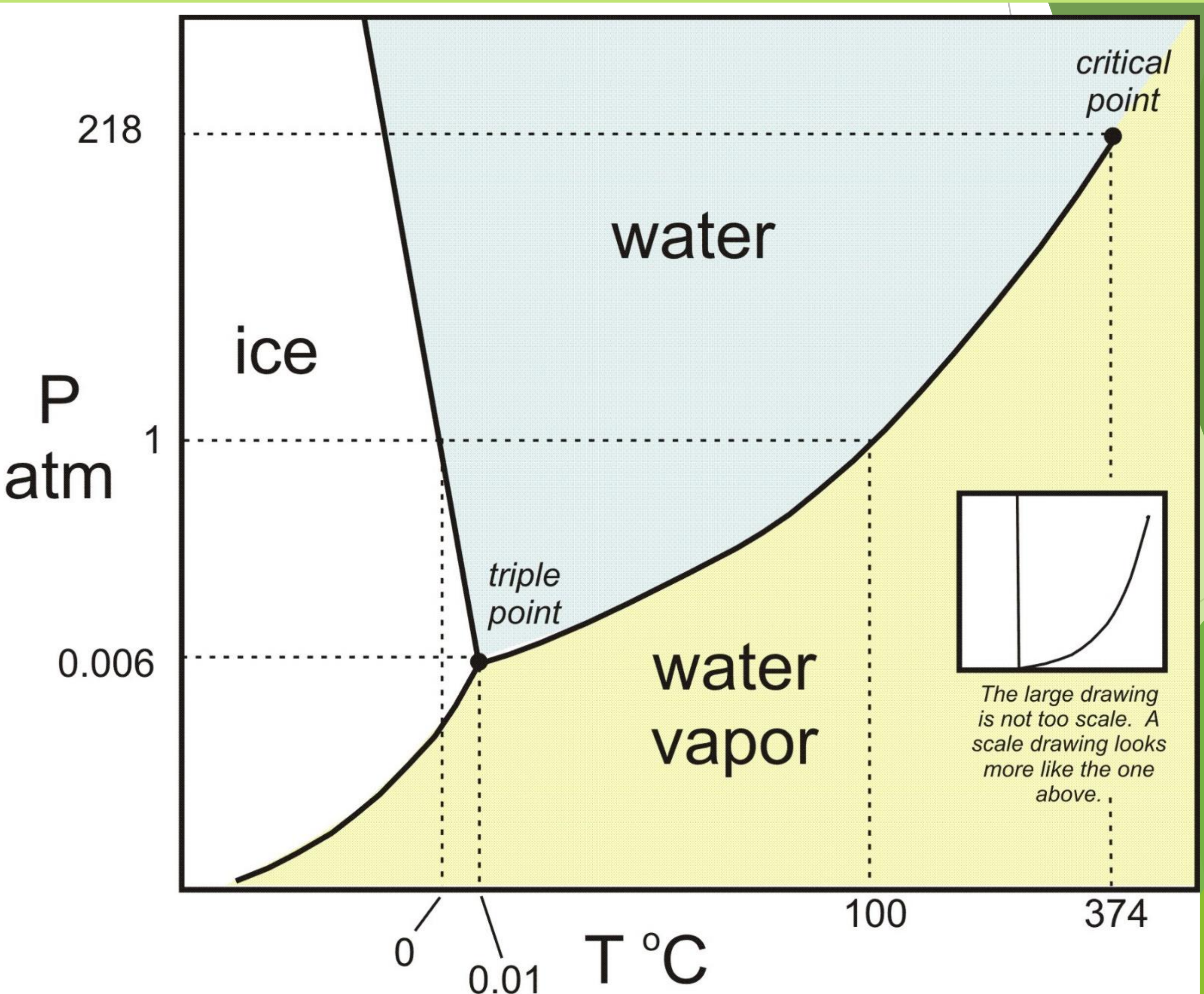
<u>Triple point-</u>

The point (temp and pressure) where solid, liquid and gas can coexist simultaneously.

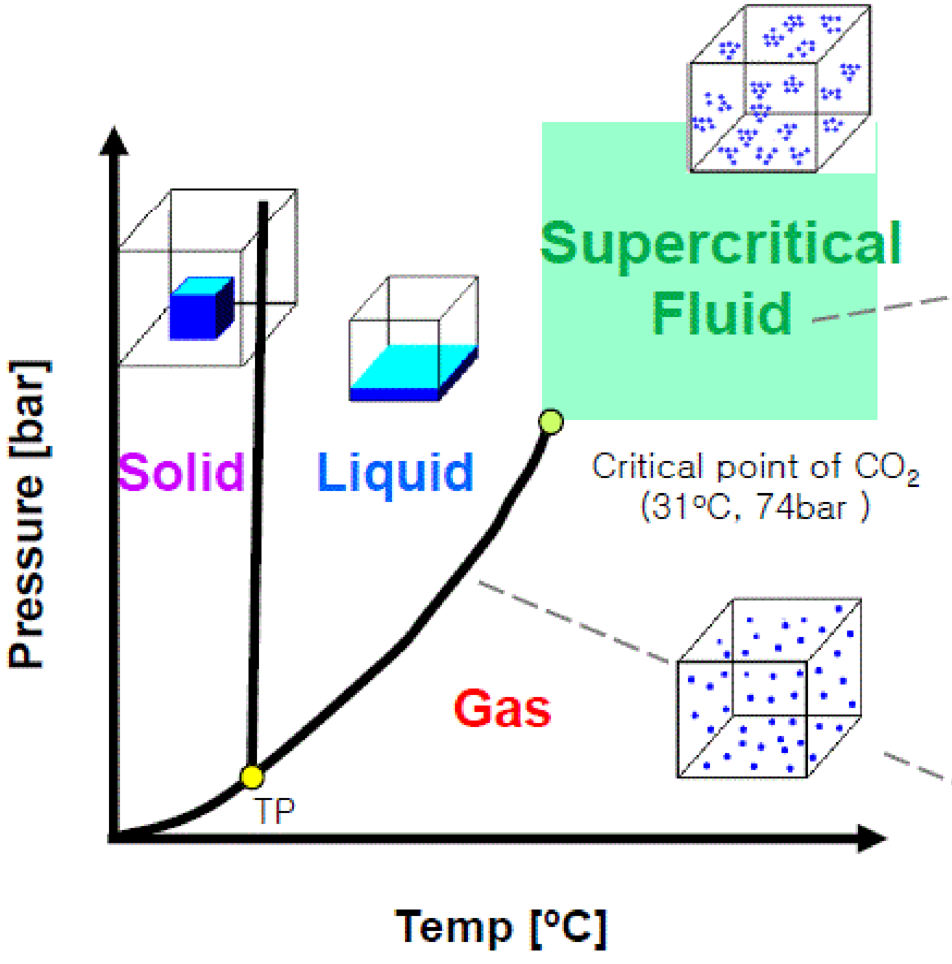


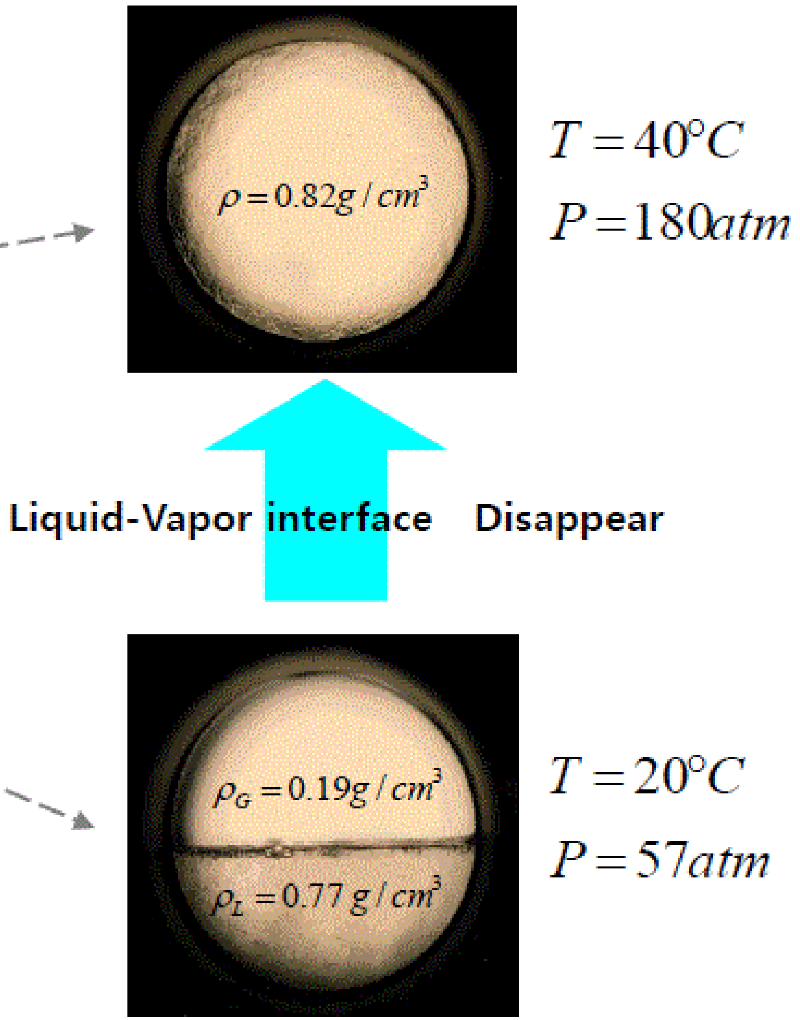
Phase Diagrams

<u>Critical point-</u>
Above this point, gas and liquid have the same densities and have odd combinations of properties and cannot be distinguished from each other



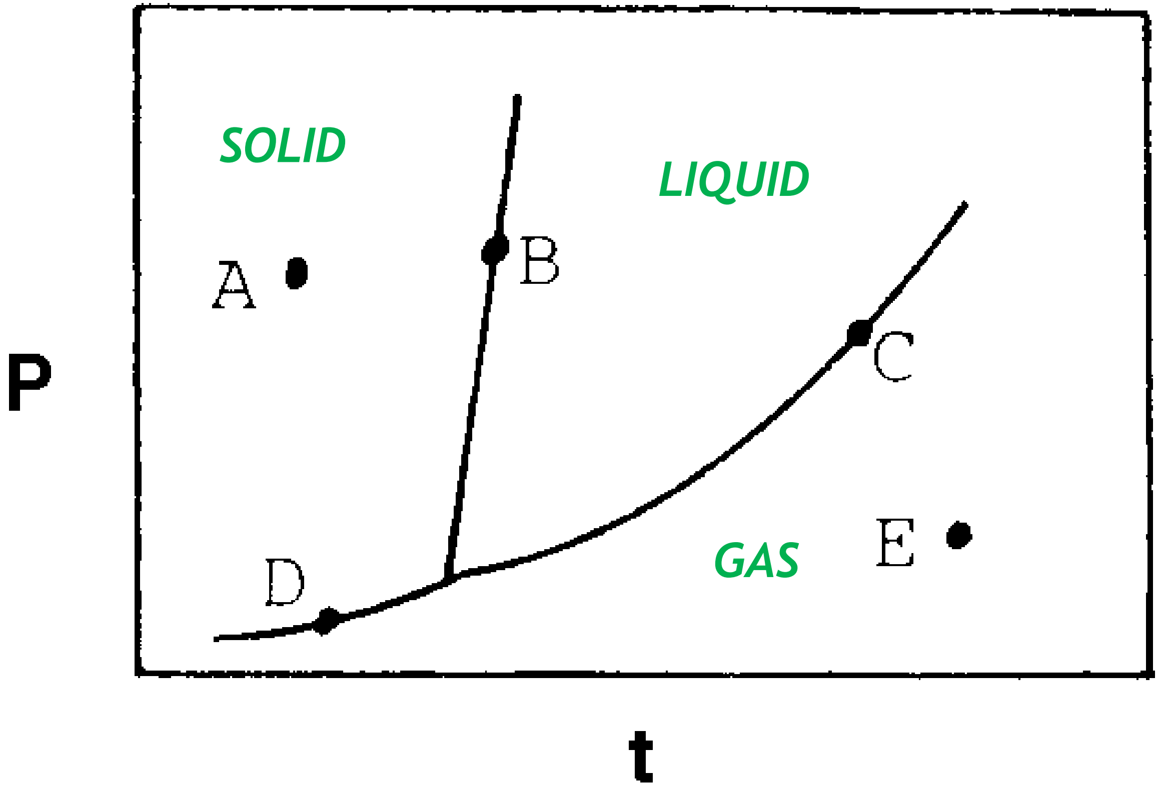
Supercritical Fluids!





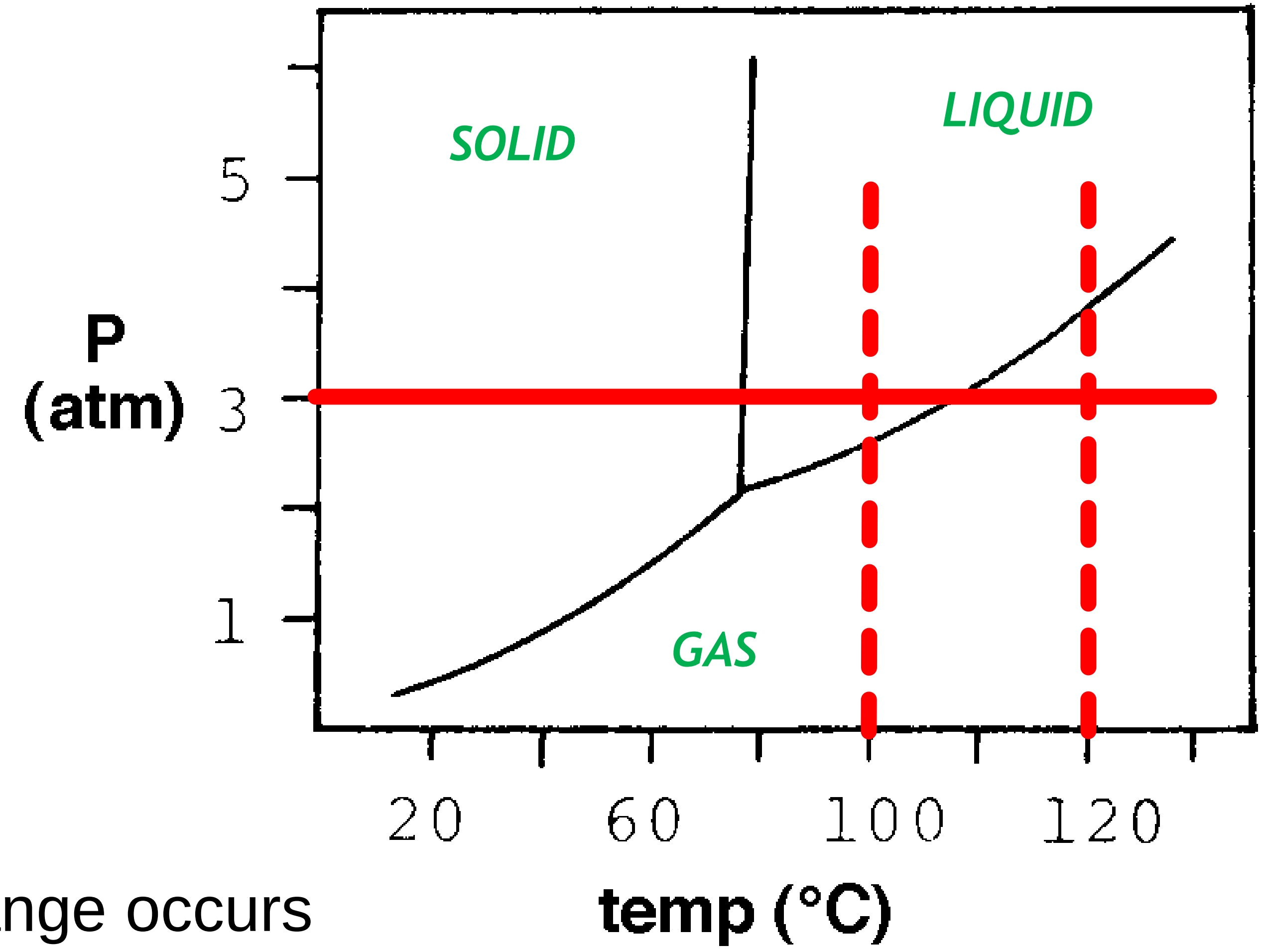
A typical phase diagram for a substance is given below. At what point on the diagram do solid and liquid exist at equilibrium?

- A
- B
- C
- D
- E



The phase diagram of a substance is given below. What occurs when the substance is heated from 100° C to 120 ° C at 3 atm pressure?

- **A** It melts
- **B** It sublimes
- **C** It vaporizes
- **D** It freezes
- **E** No phase change occurs



"Mixed Phase Calorimetry"

When you have a mixture of solids, liquids, gases

Think About It

What happens when (in terms of heat) when you add ice to your water bottle?

+

Ice is absorbing, water is releasing energy…

We can still calculate Q!

Mixtures of Solids and Liquids

<u>**The Problem:**</u> We are going to do problems that involve a phase change AND heating. Something like you drop ice into water, what temperature will the mixture be at the end?

To solve this, think about this:

- As ice melts, the temperature does not change
- BUT as soon as the ice melts the temp will rise
- We are still using

$$Q_{ice} = -Q_{water}$$

These can get tricky…

What if you have ice at -10°C and water at 50°C?

Have to heat ice AND melt it AND heat it up a bit
And also cool water down

Still have to do $Q_{ice} = -Q_{water}$ but this time…

$$(mC\Delta T + mL + mC\Delta T) = -mC\Delta T$$

Heat ice | melt ice | heat cold liq. | cool warm liq.

Practice Problem

Let's do problem #16 from WS #7 together

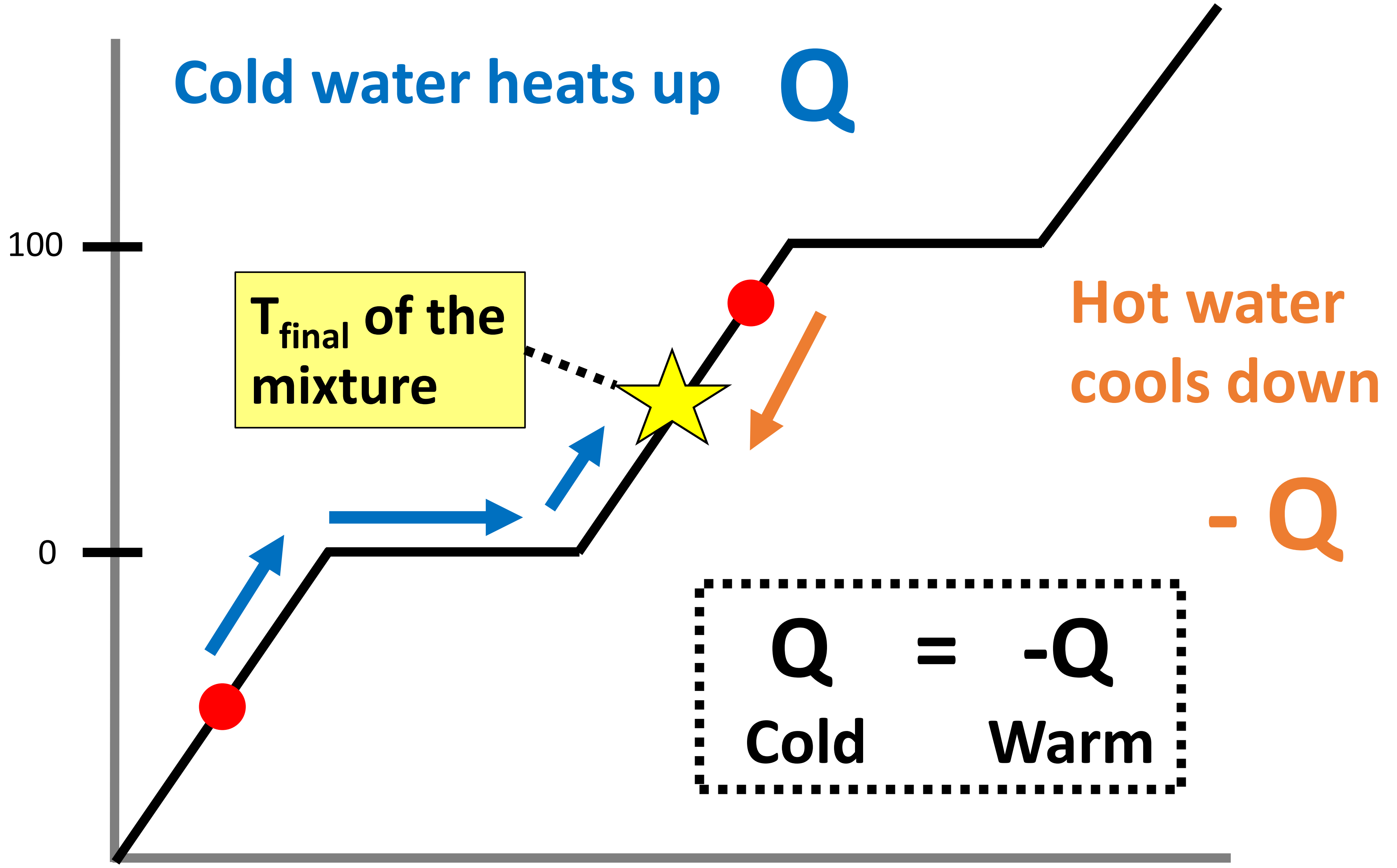
16. Determine the final temperature when 18.0 g of ice at -10.0°C mixes with 275.0 grams of water at 60.0°C

Cold

1. Heat ice
2. Melt ice
3. Heat liquid

Warm

3. Cool liquid



16. Determine the final temperature when 18.0 g of ice at -10.0°C mixes with 275.0 grams of water at 60.0°C

Qcold = -Qwarm

Cold

1. Heat ice
2. Melt ice
3. Heat liquid

Warm

3. Cool liquid

Cold

$Q_1 = mC\Delta T = (18g)(2.09 J/gC)(0° - -10°) = 376.2\ J$

$Q_2 = mL = (18g)(334\ J/g) = 6012\ J$

$Q_3 = mC\Delta T = (18g)(4.18 J/gC)(Tf° - 0°) = 75.24Tf$

$Q_{cold} = Q_1 + Q_2 + Q_3$ = 6388.2 + 75.24Tf

Warm

$Q_3 = mC\Delta T = (275g)(4.18\ J/gC)(Tf° - 60°) =$

1149.5 – 68970Tf

16. Determine the final temperature when 18.0 g of ice at -10.0°C mixes with 275.0 grams of water at 60.0°C

Qcold = -Qwarm

I personally like to do this because I don't see as many algebra mistakes or double negative issues!

Qcold = - Qwarm

Qcold + Qwarm = 0

$$6388.2 + 75.24T_f + 1149.5 - 68970T_f = 0$$

$$\frac{1224.74\,T_f}{1224.74} = \frac{62581.8}{1224.74}$$

$T_f = 51.1$ °C

Make sure your ending temperature is actually between the starting temps!

YouTube Link to Presentation

- https://youtu.be/A7z5ixKMBQs