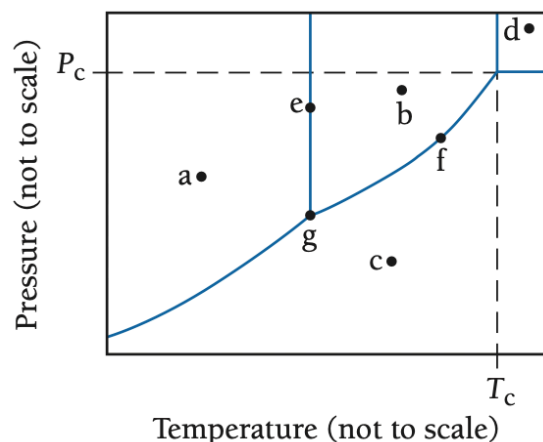


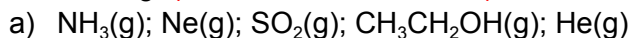
Final Review KEY

- 1) In each pair of compounds, pick the one with the higher vapor pressure at a given temperature. Explain your reasoning. (Tro 2nd Ed, Ch 11, #43)
The stronger the intermolecular forces (IMFs) of a compound are, the lower its vapor pressure and the higher its boiling point will be.
 - a) Br_2 or I_2
Both compounds are nonpolar so they will only have London dispersion forces (LDFs). Since Br is less massive than I, it will have weaker LDFs (IMFs) and thus a higher vapor pressure and lower boiling point.
 - b) H_2S or H_2O
Both compounds are polar so they will have LDFs and dipole-dipole forces (DDFs). Since H_2S doesn't exhibit hydrogen bonding like H_2O , it will have weaker IMFs and thus a higher vapor pressure and lower boiling point.
 - c) NH_3 or PH_3
Both compounds are polar so they will have LDFs and dipole-dipole forces (DDFs). Since PH_3 doesn't exhibit hydrogen bonding like NH_3 , it will have weaker IMFs and thus a higher vapor pressure and lower boiling point.
- 2) Determine if each pair of compounds forms a homogeneous solution when combined. For those that form homogeneous solutions, indicate the type of forces that are involved. (Tro 2nd Ed, Ch 11, #45)
 - a) CCl_4 and H_2O
 CCl_4 is a nonpolar molecule while H_2O is a polar molecule so the two will NOT form a homogeneous solution.
 - b) KCl and H_2O
 KCl is an ionic compound while H_2O is a polar molecule so the two WILL form a homogeneous solution due to the ion-dipole forces. LDFs, DDFs, and hydrogen bonding will also be involved.
 - c) Br_2 and CCl_4
 Br_2 is a nonpolar molecule while CCl_4 is a nonpolar molecule so the two WILL form a homogeneous solution due to the LDFs.
 - d) $\text{CH}_3\text{CH}_2\text{OH}$ and H_2O
 $\text{CH}_3\text{CH}_2\text{OH}$ is a polar molecule while H_2O is a polar molecule so the two WILL form a homogeneous solution due to the hydrogen bonding. LDFs and DDFs will also be involved.
- 3) Identify the states present at points a through g in the phase diagram shown here. (Tro 2nd Ed, Ch 11, #71)

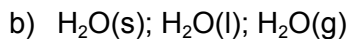


a = solid, b = liquid, c = gas, d = supercritical fluid, e = solid/liquid equilibrium, f = liquid/gas equilibrium, g = solid/liquid/gas equilibrium

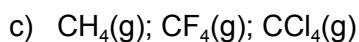
- 4) Rank each set of substances in order of increasing standard molar entropy (S°). Explain your reasoning. (Tro 2nd Ed, Ch 18, #53)



All substances are gases and atomic substances have lower entropies than molecular substances so $\text{He}(\text{g})$ and $\text{Ne}(\text{g})$ will have the lowest entropies. Further, $\text{He}(\text{g})$ has a lower molar mass than $\text{Ne}(\text{g})$ so it will have the lowest entropy and $\text{Ne}(\text{g})$ has the second lowest entropy. Meanwhile, the molecular substances will have lower entropies if they have lower molecular complexity. Therefore, $\text{SO}_2(\text{g})$ will have the middle entropy, $\text{NH}_3(\text{g})$ will have the second highest entropy and $\text{CH}_3\text{CH}_2\text{OH}(\text{g})$ will have the highest entropy.



All substances are different phases of water molecules and gases have higher entropy than liquids which have higher entropy than solids. Therefore, $\text{H}_2\text{O}(\text{s})$ will have the lowest entropy, $\text{H}_2\text{O}(\text{l})$ will have the middle entropy, and $\text{H}_2\text{O}(\text{g})$ will have the highest entropy.



All substances are gases and have the same molecular structure so they will have lower entropies if they have lower molar masses. Therefore, $\text{CH}_4(\text{g})$ will

have the lowest entropy, $\text{CF}_4(\text{g})$ will have the middle entropy, and $\text{CCl}_4(\text{g})$ will have the highest entropy.



- 5) Use data from Appendix IIB to calculate $\Delta S^\circ_{\text{rxn}}$ for each of the reactions. In each case, try to rationalize the sign of $\Delta S^\circ_{\text{rxn}}$. (Tro 2nd Ed, Ch 18, #55)
- a) $\text{C}_2\text{H}_4(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{C}_2\text{H}_6(\text{g})$
 $\Delta S^\circ_{\text{rxn}} = 1(229.2 \text{ J/K}) - 1(219.3 \text{ J/K}) - 1(130.7 \text{ J/K}) = -120.8 \text{ J/K}$
 $\Delta S^\circ_{\text{rxn}}$ is **negative** since the number of moles of gas **decreases** in the rxn
- b) $\text{C}(\text{s}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{CO}(\text{g}) + \text{H}_2(\text{g})$
 $\Delta S^\circ_{\text{rxn}} = 1(197.7 \text{ J/K}) + 1(130.7 \text{ J/K}) - 1(5.7 \text{ J/K}) - 1(188.8 \text{ J/K}) = 133.9 \text{ J/K}$
 $\Delta S^\circ_{\text{rxn}}$ is **positive** since the number of moles of gas **increases** in the rxn
- c) $\text{CO}(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2(\text{g}) + \text{CO}_2(\text{g})$
 $\Delta S^\circ_{\text{rxn}} = 1(130.7 \text{ J/K}) + 1(213.8 \text{ J/K}) - 1(197.7 \text{ J/K}) - 1(188.8 \text{ J/K}) = -42.0 \text{ J/K}$
 $\Delta S^\circ_{\text{rxn}}$ is **negative** since the molecules become **simpler** in the rxn (H_2O to H_2)
- d) $2 \text{H}_2\text{S}(\text{g}) + 3 \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{l}) + 2 \text{SO}_2(\text{g})$
 $\Delta S^\circ_{\text{rxn}} = 2(70.0 \text{ J/K}) + 2(248.2 \text{ J/K}) - 2(205.8 \text{ J/K}) - 3(205.2 \text{ J/K}) = -390.8 \text{ J/K}$
 $\Delta S^\circ_{\text{rxn}}$ is **negative** since the number of moles of gas **decreases** in the rxn
- 6) Without doing any calculations, determine the sign of ΔS_{sys} and ΔS_{surr} for each chemical reaction. (Tro 2nd Ed, Ch 18, #37)
Recall that if the entropy of the system increases then $\Delta S_{\text{sys}} > 0$ and if it decreases then $\Delta S_{\text{sys}} < 0$. Also, recall that $\Delta S_{\text{surr}} = (-\Delta H^\circ / T)$.
- a) $\text{C}_3\text{H}_8(\text{g}) + 5 \text{O}_2(\text{g}) \rightarrow 3 \text{CO}_2(\text{g}) + 4 \text{H}_2\text{O}(\text{g})$ $\Delta H^\circ_{\text{rxn}} = -2044 \text{ kJ}$
 $\Delta S_{\text{sys}} > 0$ and $\Delta S_{\text{surr}} > 0$
- b) $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{NO}(\text{g})$ $\Delta H^\circ_{\text{rxn}} = +182.6 \text{ kJ}$
 $\Delta S_{\text{sys}} < 0$ and $\Delta S_{\text{surr}} < 0$
- c) $2 \text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{N}_2\text{O}(\text{g})$ $\Delta H^\circ_{\text{rxn}} = +163.2 \text{ kJ}$
 $\Delta S_{\text{sys}} < 0$ and $\Delta S_{\text{surr}} < 0$
- d) $4 \text{NH}_3(\text{g}) + 5 \text{O}_2(\text{g}) \rightarrow 4 \text{NO}(\text{g}) + 6 \text{H}_2\text{O}(\text{g})$ $\Delta H^\circ_{\text{rxn}} = -906 \text{ kJ}$
 $\Delta S_{\text{sys}} > 0$ and $\Delta S_{\text{surr}} > 0$
- 7) Given the values of $\Delta H^\circ_{\text{rxn}}$, $\Delta S^\circ_{\text{rxn}}$, and T, determine ΔS_{univ} and predict whether each reaction is spontaneous. (Assume that all reactants and products are in their standard states.) (Tro 2nd Ed, Ch 18, #41)

- a. $\Delta H_{\text{rxn}}^{\circ} = +115 \text{ kJ}$; $\Delta S_{\text{rxn}}^{\circ} = -263 \text{ J/K}$; $T = 298 \text{ K}$
b. $\Delta H_{\text{rxn}}^{\circ} = -115 \text{ kJ}$; $\Delta S_{\text{rxn}}^{\circ} = +263 \text{ J/K}$; $T = 298 \text{ K}$
c. $\Delta H_{\text{rxn}}^{\circ} = -115 \text{ kJ}$; $\Delta S_{\text{rxn}}^{\circ} = -263 \text{ J/K}$; $T = 298 \text{ K}$
d. $\Delta H_{\text{rxn}}^{\circ} = -115 \text{ kJ}$; $\Delta S_{\text{rxn}}^{\circ} = -263 \text{ J/K}$; $T = 615 \text{ K}$

$$\Delta S_{\text{univ}} = \Delta S_{\text{system}} + \Delta S_{\text{universe}}$$

$$\Delta S_{\text{univ}} = \left(-\frac{\Delta H_{\text{sys}}}{T} \right) + \Delta S_{\text{system}}$$

$\Delta S_{\text{univ}} > 0$ for spontaneous process

a) $\Delta S_{\text{univ}} = \left(-\frac{\Delta H_{\text{sys}}}{T} \right) + \Delta S_{\text{system}}$
 $= \left(-\frac{115,000 \text{ J}}{298 \text{ K}} \right) + (-263 \text{ J/K})$
 $= -648.9 < 0$
 \rightarrow Non Spontaneous

b) Follow process a

Answer: $648.9 > 0$

\rightarrow Spontaneous

c) Follow process a

Answer: $122.81 > 0$

Spontaneous

d) Follow process a

Answer: $-76.01 < 0$

\rightarrow Non spontaneous

- 8) For each reaction, calculate $\Delta H^\circ_{\text{rxn}}$, $\Delta S^\circ_{\text{rxn}}$, and $\Delta G^\circ_{\text{rxn}}$ at 25°C and state whether the reaction is spontaneous. If the reaction is not spontaneous, would a change in temperature make it spontaneous? If so, should the temperature be raised or lowered from 25 °C? (Tro 2nd Ed, Ch 18, #61)
- $\text{N}_2\text{O}_4(\text{g}) \longrightarrow 2 \text{NO}_2(\text{g})$
 - $\text{NH}_4\text{Cl}(\text{s}) \longrightarrow \text{HCl}(\text{g}) + \text{NH}_3(\text{g})$
 - $3 \text{H}_2(\text{g}) + \text{Fe}_2\text{O}_3(\text{s}) \longrightarrow 2 \text{Fe}(\text{s}) + 3 \text{H}_2\text{O}(\text{g})$
 - $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \longrightarrow 2 \text{NH}_3(\text{g})$



$$\Delta H^\circ_{rxn} = \sum \Delta H^\circ_{f, \text{product}} - \sum \Delta H^\circ_{f, \text{reactant}}$$

$$= 2(33.2 \text{ kJ/mol}) - (9.2 \text{ kJ/mol}) = 57.2 \text{ kJ/mol}$$

$$\Delta S^\circ_{rxn} = \sum S^\circ_{f, \text{product}} - \sum S^\circ_{f, \text{reactant}}$$

$$= 2(240.1) - (304.3) = 0.175 \text{ kJ/mol}$$

$$= 175.9 \text{ J/kmol} = 0.175 \text{ kJ/mol}$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$

$$= 57.2 - (298 \text{ K})(0.175 \text{ kJ/kmol})$$

$$= 4.752 \text{ kJ/mol} \quad (\text{+ve hence non-spontaneous})$$

To be spontaneous

$$\Delta G < 0 \quad \therefore \Delta H^\circ - T\Delta S^\circ < 0$$

$$-T\Delta S^\circ < -\Delta H^\circ$$

$$T > \frac{\Delta H^\circ}{\Delta S^\circ} \Rightarrow T > \frac{57.2}{0.175} \text{ K}$$

$$\Rightarrow T > 326.85 \text{ K}$$

To make the reaction spontaneous increases the temperature above 326.85 K



$$\Delta H^\circ_{rxn} = (\Delta H^\circ_{f, HCl} + \Delta H^\circ_{f, NH_3}) - (\Delta H^\circ_{f, NH_4Cl})$$

$$= (-92 \text{ kJ/mol}) + (-46 \text{ kJ/mol}) - (-314.4 \text{ kJ/mol})$$

$$= +176.3 \text{ kJ/mol}$$

$$\Delta S^{\circ}_{\text{rxn}} = (186.7 \text{ J/Kmol}) + (192.77 \text{ J/Kmol}) - (94.85 \text{ J/Kmol})$$

$$= 284.6 \text{ J/Kmol} = 0.285 \text{ kJ/Kmol}$$

$$\Delta G^{\circ}_{\text{rxn}} = \Delta H^{\circ} - T\Delta S^{\circ}$$

$$= 176.3 - (298)(0.285) = 91.46 \text{ kJ/mol}$$

+ve hence the reaction is non-spontaneous \therefore so,
 for spontaneity $\Delta H - T\Delta S < 0$

$$T > \frac{\Delta H}{\Delta S} \Rightarrow \boxed{T > 618.6 \text{ K}}$$



$$\Delta H^{\circ}_{\text{rxn}} = (2\Delta H^{\circ}_{\text{f}}(\text{Fe}) + 3\Delta H^{\circ}_{\text{f}}(\text{H}_2\text{O})) - (3\Delta H^{\circ}_{\text{f}}(\text{H}_2) + \Delta H^{\circ}_{\text{f}}(\text{Fe}_2\text{O}_3))$$

$$= 2(0) + 3(-241.8 \text{ kJ/mol}) - (0)3 + (-822.1 \text{ kJ/mol})$$

$$= \boxed{96.7 \text{ kJ/mol}}$$

$$\Delta S^{\circ}_{\text{rxn}} = 2(0) + 3(-228.59 \text{ J/Kmol}) - (0)3 + (-741.7 \text{ J/Kmol})$$

$$\Delta S^{\circ}_{\text{rxn}} = 2(27.15) + 3(188.83) - ((130.6)3 + (90))$$

$$= 138.7 \text{ J/Kmol} =$$

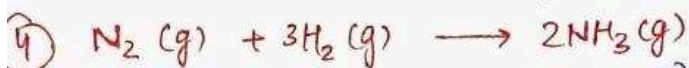
$$= 0.139 \text{ kJ/Kmol}$$

$$\Delta G^{\circ}_{\text{rxn}} = \Delta H - T\Delta S = 96.7 - (298)(0.139)$$

$$= 55.27 \text{ kJ/mol} = +ve \text{ hence non-spontaneous.}$$

$$T \Rightarrow \frac{\Delta H}{\Delta S} = \frac{96.7}{0.139} = 695.6 \text{ K}$$

To make the reaction spontaneous, increase the temp.



$$\Delta H^{\circ}_{\text{rxn}} = 2(-46.1) - (0 + 3(0)) = -92.2 \text{ kJ/mol}$$

$$\Delta S^{\circ}_{\text{rxn}} = 2(192.5) - [191.6 + (131)3] = 199.6 \text{ J/Kmol}$$

$$= 0.199 \text{ kJ/mol}$$

$$\Delta G = -92.2 + 298(0.199) = -32.9 \text{ kJ/mol}$$

-ve hence a spontaneous reaction.

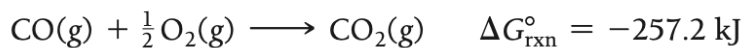
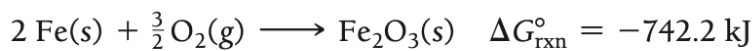
- 9) Consider the reaction:
 $2 \text{NO}(g) + \text{O}_2(g) \rightarrow 2 \text{NO}_2(g)$

Estimate ΔG° for this reaction at each temperature and predict whether the reaction is spontaneous. (Assume that ΔH° and ΔS° do not change too much within the given temperature range.) (Tro 2nd Ed, Ch 18, #65)

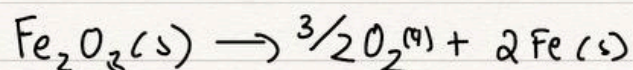
- a) 298 K
- b) 715 K
- c) 855 K

- 10) Determine ΔG° for the reaction:
 $\text{Fe}_2\text{O}_3(s) + 3 \text{CO}(g) \rightarrow 2 \text{Fe}(s) + 3 \text{CO}(g)$

Use the following reactions with known $\Delta G^\circ_{\text{rxn}}$ values: (Tro 2nd Ed, Ch 18, #67)



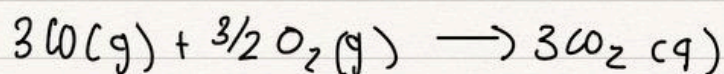
(Eq 1) Invent first Equation:



Multiply original Δb_{rxn} by -2 :

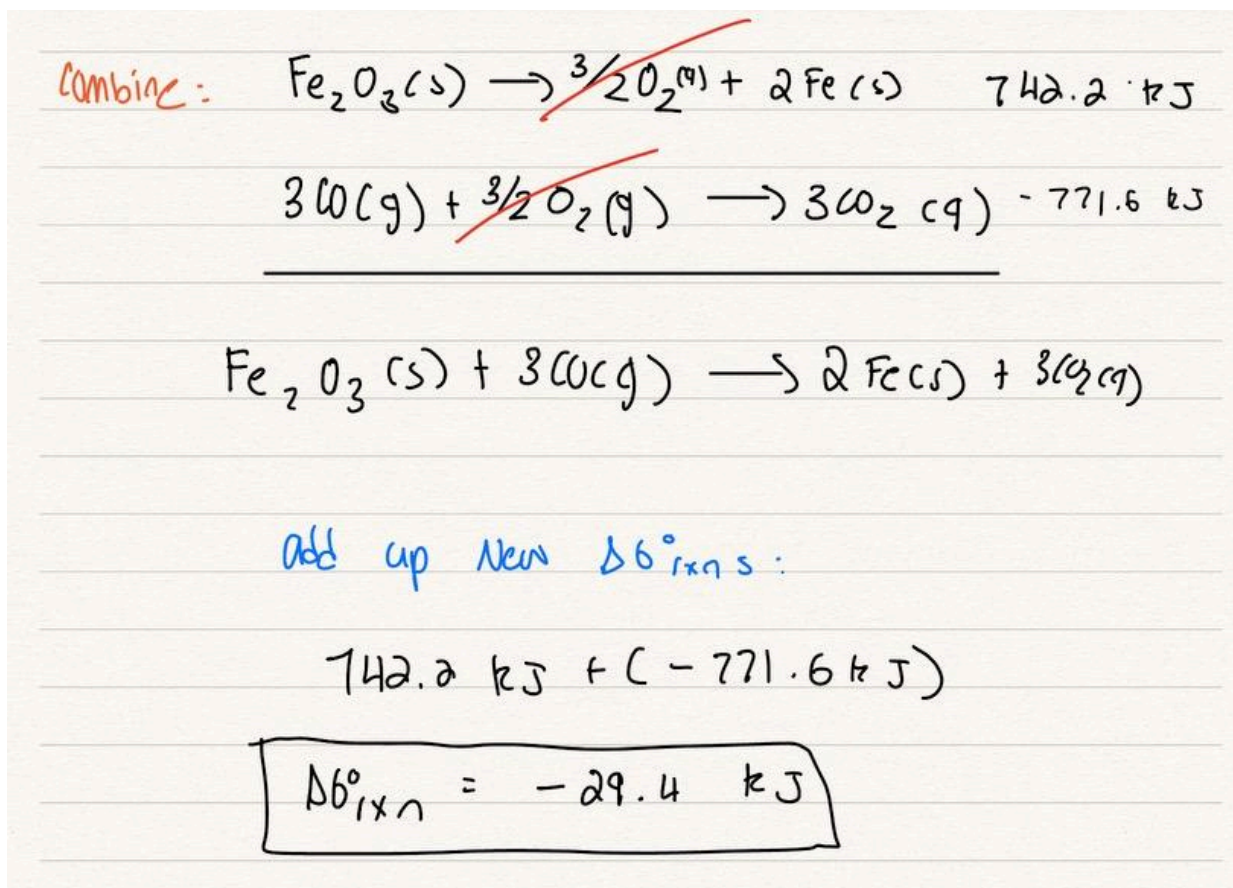
$$\begin{aligned}\Delta b_{\text{rxn}}^{\circ} &= -242.2 \text{ kJ} \cdot -2 \\ &= 742.2 \text{ kJ}\end{aligned}$$

(Eq 2) Multiply second Equation by 3



Multiply original Δb_{rxn} by 3

$$\begin{aligned}\Delta b_{\text{rxn}}^{\circ} &= -257.2 \text{ kJ} \times 3 \\ &= -771.6 \text{ kJ}\end{aligned}$$



11) Fill in the blanks in the table below (both ΔH and ΔS refer to the system). (Tro 2nd Ed, Ch 18, #47)

ΔH	ΔS	ΔG	Low Temperature	High Temperature
-	+	-	Spontaneous	_____
-	-	Temperature dependent	_____	_____
+	+	_____	_____	Spontaneous
_____	-	_____	Nonspontaneous	Nonspontaneous

ANSWER:

ΔH	ΔS	$-T\Delta S$	ΔG	Spontaneity
+	-	+	+	Nonspontaneous
-	+	-	-	Spontaneous
-	-	+	+ or -	Low Temp: Spontaneous High Temp: Nonspontaneous
+	+	-	+ or -	Low Temp: Nonspontaneous High Temp: Spontaneous

- 12) We submerge a 32.5-g iron rod, initially at 22.7 C, into an unknown mass of water at 63.2 C, in an insulated container. The final temperature of the mixture upon reaching thermal equilibrium is 59.5 C. What is the mass of the water?

When starting off these problems, it is important to identify which substance is losing heat and which substance is gaining heat. The substance that gains heat (rise in temperature) will be the one that is on the side of the equation that does not have the negative sign (left side). The iron here is gaining heat because it starts off at 22.7 deg celsius and goes to 59.6 deg celsius. The water here is on the right side with the negative sign because it LOST heat (63.2 deg celsius to 59.5 degrees celsius).

12) ① Iron (gain) water (loss)
 ↓ heat ↓ heat
 $mC\Delta T = -mC\Delta T$

② Iron Water → ③ Plug in
 $m = 32.5\text{g}$ $m = ?$
 $C = 0.451 \frac{\text{J}}{\text{g}\cdot\text{C}}$ $C = 4.186 \frac{\text{J}}{\text{g}\cdot\text{C}}$
 $T_f = 59.5^\circ\text{C}$ $T_f = 59.5^\circ\text{C}$
 $T_i = 22.7^\circ\text{C}$ $T_i = 63.2^\circ\text{C}$

④ $m_{\text{water}} = 34.72\text{g}$

- 13) We submerge a 2.85-g lead weight, initially at 10.3 C, in 7.55 g of water at 52.3 C in an insulated container. What is the final temperature of both substances at thermal equilibrium?

13.) $mC\Delta T = -mC\Delta T$

①

Lead (gained heat) Water (lost heat)

②

$m = 2.85 \text{ g}$
 $C = 0.129 \text{ J/g}^\circ\text{C}$
 $T_i = 10.3^\circ\text{C}$
 $T_f = ?$

$m = 7.55 \text{ g}$
 $C = 4.186 \text{ J/g}^\circ\text{C}$
 $T_i = 52.3^\circ\text{C}$
 $T_f = ?$

③ $(2.85 \text{ g})(0.129 \text{ J/g}^\circ\text{C})(T_f - 10.3^\circ\text{C}) = -(7.55 \text{ g})(4.186 \text{ J/g}^\circ\text{C})(T_f - 52.3^\circ\text{C})$

Isolate T_f

④ $T_f = 51.96^\circ\text{C}$

- 14) A 2.74-g sample of a substance suspected of being pure gold is warmed to 72.1 C and submerged into 15.2 g of water initially at 24.7 C. The final temperature of the mixture is 26.3 C. What is the heat capacity of the unknown substance? Could the substance be pure gold?

14.) ① $mC\Delta T = -mC\Delta T$
(water gains heat) (gold maybe lose heat)

②

Gold:	Water	mystery suspect: gold
$m = 2.74\text{g}$	$m = 15.2\text{g}$	
$C = ?$	$C = 4.186 \frac{\text{J}}{\text{g}^\circ\text{C}}$	
$T_f = 26.3^\circ\text{C}$	$T_f = 26.3^\circ\text{C}$	
$T_i = 24.7^\circ\text{C}$	$T_i = 72.1^\circ\text{C}$	

③ $(15.2\text{g})(4.186 \frac{\text{J}}{\text{g}^\circ\text{C}})(26.3^\circ\text{C} - 24.7) = -(2.74\text{g})(C)(26.3^\circ\text{C} - 72.1)$

$C_{\text{suspect}} = \underline{0.811 \frac{\text{J}}{\text{g}^\circ\text{C}}} \neq 0.129 \frac{\text{J}}{\text{g}^\circ\text{C}}$

Not Gold!

↑
literature specific heat of gold

- 15) How much heat (in kJ) is required to warm 10.0 g of ice, initially at -10.0°C , to steam at 110.0°C ? The heat capacity of ice is $2.09\text{ J/g}^{\circ}\text{C}$, and that of steam is $1.996\text{ J/g}^{\circ}\text{C}$.

15. Equations:

① $m C \Delta T = Q$ (in Joules)
 $n \Delta H_{\text{fusion}} = Q$ (in kJ); $\Delta H_{\text{fus}} = 6.01\text{ kJ/mol}$
 $n \Delta H_{\text{vap}} = Q$ (in kJ); $\Delta H_{\text{vap}} = 40.65\text{ kJ/mol}$

Water melt point = 0°C
 Water evap point = 100°C

②

③

$-10^{\circ}\text{C} \rightarrow 0^{\circ}\text{C}$ 0°C $0^{\circ}\text{C} \rightarrow 100^{\circ}\text{C}$ 100°C $100^{\circ}\text{C} \rightarrow 110^{\circ}\text{C}$

Ice + ice \rightarrow water + water + water \rightarrow gas + gas
 $m_{\text{ice}} C_{\text{ice}} \Delta T + n H_{\text{fus}} + m_{\text{water}} C_{\text{water}} \Delta T + n H_{\text{vap}} + m_{\text{steam}} C_{\text{steam}} \Delta T$
 + water & ice = different C

Water = $4.186\text{ J/g}^{\circ}\text{C}$
 Ice = $2.09\text{ J/g}^{\circ}\text{C}$

④ mass (constant throughout) + mols
 = 10.0g \rightarrow plug in! $\frac{10.0\text{g H}_2\text{O}}{18.02\text{g H}_2\text{O}} = 0.5549\text{ mols H}_2\text{O}$

$$\frac{(10.0\text{g})(4.186\frac{\text{J}}{\text{g}^{\circ}\text{C}})(0 - (-10^{\circ}\text{C}))}{1000} + (0.5549\text{ mol})(6.01\frac{\text{kJ}}{\text{mol}}) + \frac{(10.0\text{g})(4.186\frac{\text{J}}{\text{g}^{\circ}\text{C}})(100-0)}{1000} + (0.5549\text{ mol})(40.65\frac{\text{kJ}}{\text{mol}}) + \frac{(100\text{g})(1.996\frac{\text{J}}{\text{g}^{\circ}\text{C}})(110-100)}{1000}$$

$$= 30.23\text{ kJ}$$

no converts from joules to kJ

- 16) How much heat (in kJ) is evolved in converting 1.00 mol of steam at 145°C to ice at -50°C? The heat capacity of steam is 2.01 J/g°C, and that of ice is 2.09 J/g°C. The process for this will be essentially the same as the above except now you are starting in the gas state. Also remember to convert the mols to grams for the temperature change equation. In this case it will just be 1.00 mol = 18.02 grams water. The only thing to change is putting a negative sign in front of the enthalpy of fusion and vaporization equations because now we are losing heat not gaining.

16.) 145°C

gas → water
 ↓
 100°C

water → ice
 ↓
 0°C

-50°C

$n = 1.00 \text{ mol H}_2\text{O}$
 $m = 18.02 \text{ grams H}_2\text{O}$

$$\underbrace{mC\Delta T}_{\text{steam}} + \underbrace{-n\Delta H}_{\text{vap}} + \underbrace{mC\Delta T}_{\text{liquid}} + \underbrace{-\Delta H}_{\text{fusion}} + \underbrace{mC\Delta T}$$

$$\frac{(18.02\text{g})(1.996 \frac{\text{J}}{\text{g}^\circ\text{C}})(100-145^\circ\text{C})}{1000} - (1\text{mol})(40.65 \frac{\text{kJ}}{\text{mol}}) + \frac{(18.02\text{g})(4.186 \frac{\text{J}}{\text{g}^\circ\text{C}})(0-100)}{1000} - (1\text{mol})(6.01 \frac{\text{kJ}}{\text{mol}}) + \frac{(18.02\text{g})(2.09 \frac{\text{J}}{\text{g}^\circ\text{C}})(-50-0)}{1000}$$

$$\approx -57.8 \text{ kJ}$$

released!

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