Solutions to Problem Set 3

1. 1 mole liquid, T, 1 atm → vapor, T, 1 atm

Equation	Basis for the equation	Eq. #
$d\mathbf{U} = (\partial \mathbf{U}/\partial T)_V dT + (\partial \mathbf{U}/\partial V)_T dV$	$\mathbf{U} = \mathbf{U}(T,V)$	1
$\mathbf{U}(T,V)= f(T) - a/V$ $V_{liquid} = 18 \times 10^{-3}$ L mol^{-1} $V_{vapor} = 24$ L mol^{-1} $a = 5.72$ L^2 atm mol^{-1}	Given	2
$\mathbf{U}_{liquid} = f(T) - 5.72/18$ $\mathbf{U}_{vapor} = f(T) - 5.72/24$ $\mathbf{U}_{vapor} - \mathbf{U}_{liquid} = - 5.72/24 - -5.72/0.018$ $= 317.54$ L atm mol^{-1} $= 7690$ cal mol^{-1}	substituting values of V into Eq 2	3 4 5
$\mathbf{H}_{vapor} - \mathbf{H}_{liquid} = \mathbf{U}_{vapor} - \mathbf{U}_{liquid}$ $+ p(V_{vapo}r - V_{liquid})$ $= 317.54 + 1(24-0.018)$ $= 341.5$ L atm mol^{-1} $= 8270.$ cal mol^{-1} **Answer** This is smaller than $\Delta_{vap}H = 9820$ cal mol^{-1} expt	$\mathbf{H} = \mathbf{U} + pV$ Definition Using Eq 5 and values of V	

2. 0.1 mol ideal monatomic gas T_1 → T_2 along path $V = a \bullet \exp(bT)$

Equation	Basis for the equation	Eq. #
$p = 0.1RT/V$	Given ideal gas equation of state, 0.1 mol	1
$W = - \int p_{op} dV$	Definition	2
$p_{op} = p_{gas}$	given reversible	3
$W = - \int 0.1RT dV/V$		4
$V = a \bullet \exp(bT)$ $dV = a \bullet \exp(bT) \bullet b dT$ $dV/V = b dT$ $W = - \int 0.1RbT dT$ $W = - 0.1 Rb(1/2)[T_2^2 - T_1^2]$ **Answer**	Given path Taking the ln Note the function for work depends on the path!	5 6 7 8 9
$q = C (T_2 - T_1)$ However, given it is an ideal gas, for which $(\partial \mathbf{U}/\partial V)_T = 0$ $d\mathbf{U} = (\partial \mathbf{U}/\partial T)_V dT + (\partial \mathbf{U}/\partial V)_T dV$ $d\mathbf{U} = C_V dT$	Note that C is neither C_V nor C_p because the path is neither constant volume nor constant pressure. property of an ideal gas $\mathbf{U} = \mathbf{U}(T,V)$	 10 11 12

$\Delta\mathbf{U} = C_V[T_2-T_1]$ $\Delta\mathbf{U} = q + W, \quad q = \Delta\mathbf{U} - W$ $q = C_V[T_2-T_1] + (0.1Rb/2)\bullet[T_2^2-T_1^2]$ **Answer**	First law of thermodynamics	13 14 15
$T_1 = 300$ K $T_2 = 400$ K $b = 0.01$ K^{-1}	Given	16
$W = - (0.00831451/2)(400^2-300^2)$ $W = - 291$ J $= -0.291$ kJ **Answer**	Using Eq 9 and 0.1 mol	17 18
$\Delta\mathbf{U} = 1.5(8.31451)[400-300] = 1247$ J $= 1.247$ kJ **Answer**		19 20
$q = \Delta\mathbf{U} - W = 1.538$ J **Answer**	Using Eq 14, 18 and 20	

3. **A** **B** **C**

1 mole ideal gas $p_1, T_1, V_1 \rightarrow p_2, T_2, V_2 \rightarrow p_3, T_3, V_3 \rightarrow p_4, T_4, V_4$

A reversible heating at constant p = 1 atm, from 298 to 373 K

B reversible isothermal expansion to $2V_2$

C reversible adiabatic cooling to 308 K

$C_p = (5/2)R$

Equation	Basis for the equation	Eq. #
$p_1 = 1$ atm $T_1 = 298$ $V_1 = R298/1$ atm $= 24.45$ L	Given ideal gas	1 2
$p_2 = 1$ atm $T_2 = 373$ $V_2 = R373/1$ atm $= 30.61$ L		3 4
$T_3 = T_2 = 373$ $V_3 = 2V_2 = 61.22$ L $p_3 = R373/61.22 = 0.50$ atm		5 6
$T_4 = 308$ K p_4 , V_4 can be calculated but we actually do not need these		7
$d\mathbf{U} = C_V dT + (\partial\mathbf{U}/\partial V)_T dV$	$\mathbf{U} = \mathbf{U}(T,V)$ and definition of C_V	8
$(\partial\mathbf{U}/\partial V)_T = 0$ $d\mathbf{U} = C_V dT$	property of an ideal gas	9 10
$d\mathbf{H} = C_p dT + (\partial\mathbf{H}/\partial p)_T dp$	$\mathbf{H} = \mathbf{H}(T,p)$ and definition of C_p	11
$(\partial\mathbf{H}/\partial p)_T = 0$ $d\mathbf{H} = C_p dT$	property of an ideal gas	12 13
Step A $\Delta\mathbf{U} = (3/2)R(373-298) = 935.4$ J **Answer**	From Eq 10 and given $C_V = (3/2)R$	14
$\Delta\mathbf{H} = (5/2)R(373-298) = 1559.0$ J **Answer**	From Eq 13 and given $C_p = C_V + R = (5/2)R$	15
$W = - \int p_{op} dV = -1$ atm $[V_2-V_1]$ $= -1\,[30.61 - 24.45] = - 6.16$ L atm $= - 624.2$ J **Answer**	Definition and p = constant = 1 atm	16 17

$\Delta\mathbf{U} = q + W$ $q = \Delta\mathbf{U} - W = 935.4 + 624.2 = 1559.6$ J **Answer**	First law of thermodynamics	18 19
Step B $\Delta\mathbf{U} = 0$ J **Answer**	$dT = 0$	20
$\Delta\mathbf{H} = 0$ J **Answer**	$dT = 0$	21
$W = - \int p_{op}dV = - \int p_{ga}sdV$ $W = - \int RTdV/V = - R373 \ln(V_3/V_2)$ $W = - R373 \ln 2 = -2149$ J **Answer**	$p_{op} = p_{gas}$ reversible expansion	22 23 24
$q = \Delta\mathbf{U} - W$ $q = 0 + 2149$ J **Answer**	First law of thermodynamics	25 26
Step C $\Delta\mathbf{U} = (3/2)R(308-373) = - 810.7$ J **Answer**	From Eq 10 and given $_{CV} = (3/2)R$	27
$\Delta\mathbf{H} = (5/2)R(308-373) = - 1351.1$ J **Answer**	From Eq 13 and given $C_p=C_V+R = (5/2)R$	
$q = 0$ **Answer**	adiabatic	
$\Delta\mathbf{U} = q + W$ $W = \Delta\mathbf{U}$ $W = - 810.7$ J **Answer**	First law of thermodynamics From Eq 27	
Overall $\Delta\mathbf{U} = 124.7$ J $\Delta\mathbf{H} = 208.8$ J $W = -3583.9$ $q = 3708.6$ J **Answer**		

4. **A** **B** **C**

1 mole ideal gas $p_1, T_1, V_1 \rightarrow p_2, T_2, V_2 \rightarrow p_3, T_3, V_3 \rightarrow p_4, T_4, V_4$

A reversible heating at constant V, from 298 to 373 K

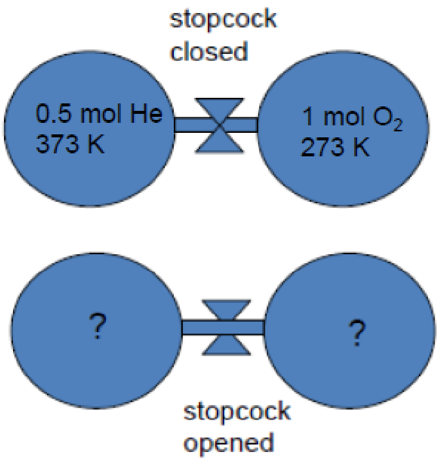
B Joule expansion into a vacuum to $2V_2$

C reversible cooling at constant p to 298 K

Equation	Basis for the equation	Eq. #
$p_1 = ?$ $T_1 = 298$ $V_1 = ?$	Given	
$p_2 = ?$ $T_2 = 373$ $V_2 = V_1$		
$p_3 = ?$ $T_3 = 373$ $V_3 = 2V_2$	In a Joule expansion, there is no T change	
$p_4 = p_3$ $T_4 = 298$ K $V_4 = ?$		
Step A $q = C_V(373-298)$ $W = - \int p_opdV = 0$ since $dV = 0$		 1 2
$d\mathbf{U} = C_VdT + (\partial\mathbf{U}/\partial V)_TdV$ $(\partial\mathbf{U}/\partial V)_T = 0$ $\Delta\mathbf{U} = C_V(373-298) = q + W$ indeed	$\mathbf{U} = \mathbf{U}(T,V)$ and definition of C_V property of an ideal gas	3 4 5

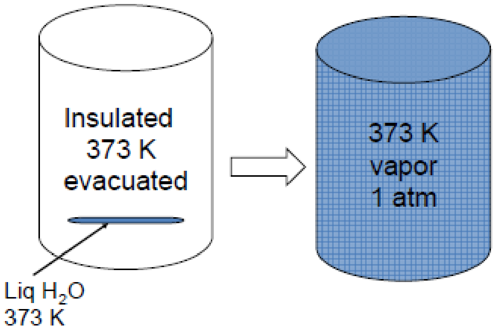
$d\mathbf{H} = C_p dT + (\partial\mathbf{H}/\partial p)_T dp$ $(\partial\mathbf{H}/\partial p)_T = 0$ $\Delta\mathbf{H} = C_p(373-298)$ Another way, $\Delta\mathbf{H} = \Delta\mathbf{U} + p_2V_2 - p_1V_1$ $= C_V(373-298) + R373 - R298$ $= C_p(373-298)$ the same	**H** = **H**(T,p) and definition of C_p property of an ideal gas **H** = **U** + pV definition Using ideal gas law $C_p - C_V = R$ for an ideal gas	6 7 8 9 10 11
Step B $W = -\int p_{op}dV$ $p_{op} = 0$ $W = 0$ $\Delta\mathbf{U} = C_V dT = 0$ $\Delta\mathbf{U} = q + W$ $q = \Delta\mathbf{U} = 0$ $\Delta\mathbf{H} = C_p dT = 0$	dT = 0 in a Joule expansion	12 13 14 15 16 17 18
Step C $q = C_p(298-373)$ $\Delta\mathbf{U} = C_V(298-373)$ $\Delta\mathbf{U} = q + W$ $W = C_V(298-373) - C_p(298-373)$ $= -R(298 - 373)$ Another way: $W = -\int p_{op}dV$ reversible pop = p_{gas} At constant pressure, $W = -p_3\int dV = -p_3(V_4-V_3) = -R298 - R373$ the same as $-R(298 - 373)$ found from first law $\Delta\mathbf{H} = C_p(298 - 373)$	.	19 20 21 22 23 24
for the overall process $q = C_V(373-298) + 0 + C_p(298-373)$ $= (C_p-C_V)298 - (C_p-C_V)373$ $q = R(298-373) = -0.624$ kJ **Answer** $W = 0+0 - R(298 - 373) = +0.624$ kJ **Answer** $\Delta\mathbf{U} = C_V(373-298) + 0 + C_V(298-373)$ $\Delta\mathbf{U} = 0$ **Answer** $\Delta\mathbf{H} = C_p(373-298) + 0 + C_p(298-373)$ $\Delta\mathbf{H} = 0$ **Answer**	$C_p-C_V = R$ for an ideal gas Note answers do not require C_p or C_V values.	

5. adiabatic process, ideal gases



Equation	Basis for the equation	Eq. #
$q=0$	Given adiabatic	1
$W = 0$	No work done to surroundings	2
$\Delta \mathbf{U} = q + W = 0$		3
$d\mathbf{U} = C_V dT + (\partial \mathbf{U}/\partial V)_T dV$ $(\partial \mathbf{U}/\partial V)_T = 0$ $\Delta \mathbf{U} = C_V(T_f - T_i)$ $\Delta \mathbf{U}_{He} = 0.5\,(3/2)R(T_f - 373)$ $\Delta \mathbf{U}_{O2} = (5/2)R(T_f - 273)$	$\mathbf{U} = \mathbf{U}(T,V)$ and Definition of C_V property of an ideal gas	4 5 6 7 8
$0 = \Delta \mathbf{U}_{He} + \Delta \mathbf{U}_{O2}$ $0 = 0.5(3/2)R(T_f - 373) + (5/2)R(T_f - 273)$ $T_f = \frac{(3/4)373 + (5/2)273}{(3/4)+(5/2)}$ $T_f = 296$ K **Answer**	From Eq 3 Solving for T_f	9

6. 10 L tank, ideal gas vapor, negligible volume of liquid water



Equation	Basis for the equation	Eq. #
$q=0$	Given insulated, no q to or from surroundings	1
$W = 0$	vapor formed in evacuated volume	2
Final state: water vapor $V = 10$ L, $T = 373$, $p = 1$ atm in equilibrium with remaining liquid		
$pV = nRT$	ideal gas	3
1 atm (10 L) $= n\ 0.0820578$ L atm mol^{-1} K^{-1} (373)		
$n = 0.3267$ mol		4
$\Delta \mathbf{H} = 9720$ cal mol^{-1}(0.3267 mol)	given $\Delta_{vap}\mathbf{H}$ and using moles from Eq 4	
$= 3175$ cal **Answer**		5
$\Delta \mathbf{U} = q + W + \mathbf{U}_{vap} - \mathbf{U}_{liq} = \Delta_{vap}\mathbf{U}$	first law	6
$\Delta \mathbf{H} = \Delta \mathbf{U} + (pV)_{vap} - (pV)_{liq}$	**H** = **U** + pV definition	7
$\Delta \mathbf{H} = \Delta_{vap}\mathbf{U} + 1\text{atm}(10\text{ L}) - (\sim 0)$	neglecting V_{liq} remaining in final state and neglecting $(pV)_{liq}$ at the initial state	8
3175 cal $= \Delta_{vap}\mathbf{U}$ $+ 1\text{atm}(10\text{ L})(1.9872/.0820578)$	substituting value from Eq 5 into Eq 8 converting to cal from L atm	
$\Delta_{vap}\mathbf{U} = 3175 - 242.2 = 2933$ cal **Answer**		

7. (a) 10 g ice(263 K) + 25 g liq water (288 K) → **?** adiabatic process

Equation	Basis for the equation	Eq. #
(i) 10 g ice(263 K) → 10 g ice (273 K) (ii) 10 g ice(273 K) → 10 g liq (273 K) (iiii) 10 g liq(273 K) → 10 g liq (T_f K) (iv) 25 g liq (288 K) → 25 g liq (T_f K)	One possibility is that all ice melts and final state is all liq water at $T_f \geq 273$. Another possibility is that only some of the ice melts and final state is some ice, some liquid water, all at 273 K The largest contribution comes from the q associated with melting of ice, so the second possibility is more likely. Nevertheless, let us try the first possibility and see what we find out.	
$q = 0 = q_i + q_{ii} + q_{iii} + q_{iv}$	q contributions must sum to zero because it is an adiabatic process.	
$q_i = 10\text{g}\bullet 0.5 \text{ cal K}^{-1}\text{g}^{-1}\bullet(273-263)\text{K}$	$q = \int C dT$, and C is independent of temperature	
$q_{ii} = 10 \text{ g}\bullet 80 \text{ cal g}^{-1}$ $q_{iii} = 10 \text{ g}\bullet 1.0 \text{ cal K}^{-1}\text{g}^{-1}\bullet(T_f-273)\text{K}$ $q_{iv} = 25 \text{ g}\bullet 1.0 \text{ cal K}^{-1}\text{g}^{-1}\bullet(T_f-288)\text{K}$ $0 = 50 + 800 + 10T_f - 2730 + 25T_f - 7200$ Solving for T_f		
$T_f = 259.5$ K this answer is **absurd**.	because it is lower than the freezing point of water. Therefore it is not possible that all ice melts	

(i) 10 g ice(263 K) → 10 g ice (273 K) (ii) x g ice(273 K) → x g liq (273 K) (iiii) 25 g liq(288 K) → 25 g liq (273 K)	Assume only some of the ice melts and the final state is some ice some liquid all at 273 K	
$q = 0 = q_i + q_{ii} + q_{iii}$ $q_i = 10 g \bullet 0.5$ cal $K^{-1} g^{-1} \bullet (273-263)$K $q_{ii} = x\ g \bullet 80$ cal g^{-1} $q_{iii} = 25\ g \bullet 1.0$ cal $K^{-1} g^{-1} \bullet (273-288)$K	q contributions must sum to zero because it is an adiabatic process	
$0 = 50 + 80x – 375$ x = 4.06 g of ice melted Final state is 5.94 g ice 29.06 g liq water all at 273 K **Answer**	solve for x	

7. (b) 10 g ice(263 K) + 100 g liq water (288 K) → **?** adiabatic process

Equation	Basis for the equation	Eq. #
(i) 10 g ice(263 K) → 10 g ice (273 K)	One possibility is that all ice melts and final state is all liq water at $T_f \geq 273$. Another possibility is that only some of the ice melts and final state is some ice, some liquid water, all at 273 K Let us try the first possibility and see what we find out.	1
(ii) 10 g ice(273 K) → 10 g liq (273 K)		2
(iiii) 10 g liq(273 K) → 10 g liq (T_f K)		3
(iv) 100 g liq (288 K) → 100 g liq (T_f K)		4
$q = 0 = q_i + q_{ii} + q_{iii} + q_{iv}$	Adiabatic process	5
$q_i = 10 g \bullet 0.5$ cal $K^{-1} g^{-1} \bullet (273-263)$K		6
$q_{ii} = 10\ g \bullet 80$ cal g^{-1}		7
$q_{iii} = 10\ g \bullet 1.0$ cal $K^{-1} g^{-1} \bullet (T_f-273)$K		8
$q_{iv} = 100\ g \bullet 1.0$ cal $K^{-1} g^{-1} \bullet (T_f-288)$K		9
$0 = 50 + 800 + 10T_f - 2730 + 100T_f – 28800$	Substituting each term into Eq 5	10
Solving for T_f $T_f = 278.9$ K The final state is 110 g liquid water at 278.9 K **Answer**		

8. C + $H_2O(g) \rightarrow CO(g) + H_2(g)$ rxn (1)
$CO(g) + H_2O(g) \rightarrow CO_2(g) + H_2(g)$ rxn (2)

Equation	Basis for the equation	Eq. #
(a) Final products are equimolar $CO(g) + H_2(g)$	Assuming only rxn (1) occurs	1
Combustion at 298 K, 1 atm:	Given	
$CO(g) + H_2(g) + O2 \rightarrow CO_2(g) + H_2O(g)$	Assuming all combustion products are vapor	2
$\Delta_{rxn}\mathbf{H} = \Delta_{form}\mathbf{H}[CO_2(g)] + \Delta_{form}\mathbf{H}[H_2O(g)] - \Delta_{form}\mathbf{H}[CO(g)]$	By definition, $\Delta_{form}\mathbf{H}[H_2(g)] = 0$ $\Delta_{form}\mathbf{H}[O_2(g)] = 0$	3

$\Delta_{rxn}\mathbf{H}$ = -393.51 -241.82 –(-110.53) = -524.8 kJ mol^{-1}	Look up $\Delta_{form}\mathbf{H}$ for $CO_2(g)$, $H_2O(g)$ and CO(g) in textbook at 1 atm, 298 K in kJ mol^{-1}	4
For 1 L of water gas pV = nRT n =1 atm•1 L /(0.0820578L atm mol^{-1} K^{-1} •298) n = 0.0409 mol = 0.02045 mol CO + 0.02045 mol H_2 $\Delta_{comb}\mathbf{H}$ of 1 L of water gas = 0.02045 mol•(-524.8 kJ mol^{-1}) = -10.73 kJ **Answer**	Assume ideal behavior Ideal gas law Solve for total number of moles in 1 L of gas for this number of moles	5 6 7 8
(b) Final products of water gas reaction are 0.95 CO(g) +0.05 $CO_2(g)$ +1.05$H_2(g)$ For 1 L of water gas, n = 0.0409 mol as before, = 0.01895 mol CO +0.0010 mol CO_2 + 0.02095 mol H_2 $CO(g) + ½O_2 \rightarrow CO_2(g)$ $\Delta_{rxn}\mathbf{H}$ = $\Delta_{form}\mathbf{H}[CO_2(g)]$ - $\Delta_{form}\mathbf{H}[CO(g)]$ = -393.51–(-110.53) = -282.98 kJ mol^{-1} $H_2(g) + (1/2)O_2 \rightarrow H_2O(g)$ $\Delta_{rxn}\mathbf{H}$ = $\Delta_{form}\mathbf{H}[H_2O(g)]$ = -241.82 kJ mol^{-1} $\Delta_{comb}\mathbf{H}$ of 1 L of water gas = 0.01895 mol • -282.98 kJ mol^{-1} + 0.02095 mol • -241.82 kJ mol^{-1} = -10.42 kJ **Answer**	95% of C appears as CO from rxn (1) and and 5% as CO_2 from rxn (2) but now consists of (0.95/2.05)(0.0409) mol CO + (0.05/2.05)(0.0409) mol CO_2 + (1.05/2.05)(0.0409) mol H_2 Do $\Delta_{rxn}\mathbf{H}$ separately because no longer 1:1 in CO and H_2 Using $\Delta_{form}\mathbf{H}$ for $CO_2(g)$ and CO(g) Using $\Delta_{form}\mathbf{H}$ for $H_2O(g)$ Using the number of moles found in Eq 10 Enthalpy of combustion of 1 L of water gas	9 10 11 12

9. $C + ½O_2(g) \rightarrow CO(g)$ rxn (1)
$CO(g) + ½O_2(g) \rightarrow CO_2(g)$ rxn (2)

Equation	Basis for the equation	Eq. #
Final products are 0.95 mol CO + 0.05 mol CO_2 for every 0.5 mol O_2 used up, which is accompanied by 2.0 mol N_2 in the dry air used to produce the gas	95% of O_2 appears in CO from rxn (1) and and 5% in CO_2 from rxn (2)	
consider 1 L producer gas at 1 atm pV = nRT n =1 atm•1 L /(0.0820578L atm mol^{-1} K^{-1} •298) n = 0.0409 mol = (0.95/3.0)(0.0409) mol CO + .	Ideal gas law	1 2 3

(0.05/3.0)(0.0409 mol CO_2 + (2.0/3.0)(0.0409) mol N_2 or 0.01295 mol CO	Of the components only CO can contribute to heating value	
Combustion of CO in air $CO(g) + ½O_2(g) \rightarrow CO_2(g)$ $\Delta_{rxn}\mathbf{H} = \Delta_{form}\mathbf{H}[CO_2(g)] - \Delta_{form}\mathbf{H}[CO(g)]$ = -393.51–(-110.53) = -282.98 kJ mol^{-1}	By definition, $\Delta_{form}\mathbf{H}[O_2(g)] = 0$ using $\Delta_{form}\mathbf{H}$ for $CO_2(g)$ and CO(g) in textbook at 1 atm, 298 K in kJ mol^{-1}	4 5 6
$\Delta_{comb}\mathbf{H}$ = 0.01295 mol • -282.98 kJ mol^{-1} = 3.66 kJ per L of producer gas **Answer**	For 1 L producer gas containing 0.01295 mol CO	7

10. (a) $CH_4(g) + 2O_2(g)$/air at 293 K → $CO_2(g) + 2H_2O(g)$ at 373 K

Equation	Basis for the equation	Eq. #
1 mol CH_4 4 mol O_2 16 mol N_2	given Twice the minimum amount of air is used for combustion	
At 293 K $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$ $\Delta_{comb}\mathbf{H}$ = -212.91 kcal mol^{-1}	Definition of $\Delta_{comb}\mathbf{H}$ is that products of combustion are $CO_2(g)$ and $H_2O(l)$ Given	1 2
$CO_2(g)$ 293 K → $CO_2(g)$ 373 K (1) 2H2O (l) 293 K →2H2O (l) 373 K (2) 2H2O (l) 373 K →2H2O (g) 373 K (3) $16N_2(g)$ 293 K →$16N_2(g)$ 373 K (4) $2O_2(g)$ 293 K → $2O_2(g)$ 373 K (5) $CH_4(g) + 4O_2(g)$ +16N2(g) at 293 K → $CO_2(g) + 2H_2O(g) + 2O_2 + 16N_2$ at 373 K	The CO2 product has to be heated to 373 K the liquid water has to be heated to 373 K and vaporized The N_2 and the unused O_2 have to be heated to 373 K Adding up the changes	3 4 5 6 7 8
$\Delta\mathbf{H} = \Delta_{comb}\mathbf{H} + \Delta\mathbf{H}(1) + \Delta\mathbf{H}(2) + \Delta\mathbf{H}(3) + \Delta\mathbf{H}(4) + \Delta\mathbf{H}(5)$ $\Delta\mathbf{H}(1)$ = 1•8.87 cal mol^{-1} K^{-1} (373-293) $\Delta\mathbf{H}(2)$ =2•18 cal mol^{-1} K^{-1} (373-293) $\Delta\mathbf{H}(3)$ =2•9720 cal mol^{-1} $\Delta\mathbf{H}(4)$ =16•6.96 cal mol^{-1} K^{-1} • (373-293) $\Delta\mathbf{H}(5)$ =2•7.02 cal mol^{-1} K^{-1} (373-293)	Using d$\mathbf{H}$ = C_pdT at dp =0 Using $\Delta_{vap}\mathbf{H}$ =9720 cal mol^{-1} for water	9 10 11 12 13 14
$\Delta\mathbf{H}$ = -212.91×10^3 + 709.6 + 2880 +19440 + 8909 + 1123 = -179848 cal **Answer**	Summing up terms in Eq 9 from Eq 10-14 heat liberated for 1 mol of CH_4 gas burned	

10. (b) $CH_4(g) + 2O_2(g)$/air at 293 K → $CO_2(g) + 2H_2O(g)$ at T K

Equation	Basis for the equation	Eq. #
1 mol CH_4 4 mol O_2 16 mol N_2	given Twice the minimum amount of air is used for combustion	
final state of stack gases: T such that saturated with water vapor but no condensation Composition of stack gases: 2 mol O_2 +16 mol N_2 +1 mol CO_2 + 2 mol H_2O = 21 mol $x_{H2O} = 2/21 = 0.095$ 0.095 atm is the partial pressure of water at 1 atm total pressure or 72 mm Hg	given find the mole fraction of water in the stack gases using plot of vapor pressure of water, 72 mm Hg corresponds to 40°C or 313 K	2 3
same steps Eq 3-8 as in 10.(a) have to be added up except T is 313 not 373	Everything is the same as in **10. (a)** except that the temperature of the stack gases is 313 K instead of 373 K and we need to get $H_2O(g)$ from 373 K back down to 313 K	
$\Delta\mathbf{H} = \Delta_{comb}\mathbf{H} + \Delta\mathbf{H}(1) + \Delta\mathbf{H}(2) + \Delta\mathbf{H}(3) + \Delta\mathbf{H}(2a) + \Delta\mathbf{H}(4) + \Delta\mathbf{H}(5)$ $\Delta\mathbf{H}(1) = 1\bullet 8.87$ cal mol^{-1} K^{-1} (313-293) $\Delta\mathbf{H}(2) = 2\bullet 18$ cal mol^{-1} K^{-1} (373-293) $\Delta\mathbf{H}(3) = 2\bullet 9720$ cal mol^{-1} $\Delta\mathbf{H}(2a) = 2\bullet 8.03$ cal mol^{-1} K^{-1} (313-373) $\Delta\mathbf{H}(4) = 16\bullet 6.96$ cal mol^{-1} K^{-1} $\bullet$ (313-293) $\Delta\mathbf{H}(5) = 2\bullet 7.02$ cal mol^{-1} K^{-1} (313-293)	Using d**H** = C_pdT at dp =0 Using $\Delta_{vap}\mathbf{H}$ =9720 cal mol^{-1} for water at 373 K	4 5 6 7 8 9 10
$\Delta\mathbf{H} = -212.91\times10^3 + 177 + 2880 + 19440 - 321 + 2227 + 280$ = -188227 cal **Answer**	Summing up Eq 5-10 into Eq 4 heat liberated for 1 mol of CH_4 gas burned to stack gases at 313 K where water in saturated vapor just fails to condense	